Redox reaction

Redox:

- Those chemical reactions, in which reduction and oxidation both are carried out simultaneously are redox reaction.

Oxidation and Reduction:

★	OXIDATION	REDUCTION
	- loss of electron eg: $Fe \longrightarrow Fe^{2+} + 2e^-$	- gain of electron eg: $Fe^{2+} + 2e^- \longrightarrow Fe$.
	- addition of oxygen eg: $CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$	- removal of oxygen eg: $CaCO_3 \longrightarrow CaO + CO$ or $CO_2 \longrightarrow C + O_2$.
	- removal of hydrogen eg: $2NH_3 \longrightarrow N_2 + 3H_2$	- addition of hydrogen eg: $N_2 + 3H_2 \longrightarrow 2NH_3$
	- oxidation number increases. (+↑) (−↓)	- oxidation number decreases. (+↓) (−↑)

Rules to find Oxidation no.:

Fluorine ⟶ −1	Hydrogen ⟶ −1 (metal hydride)
IA Group ⟶ +1	Oxygen ⟶ −2 (98%)
IIA Group ⟶ +2	Oxygen ⟶ −1 (H_2O_2) / −½ (KO_2)
Hydrogen ⟶ +1 (99% time)	Free state element ⟶ 0.

Sum of oxidation state → 0 in individual molecule.	Sum of oxidation state → charge on ion in ion

Q:1 Find oxidation number, renderlined atom:

(a) $\underline{CH_4}$ (b) $\underline{SO_2}$ (c) $\underline{CaH_2}$ (d) $\underline{OF_2}$

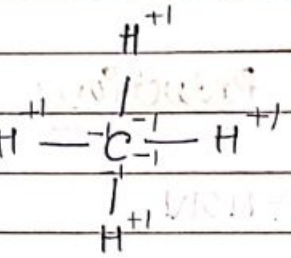
(e) $\underline{KMnO_4}$ (f) $\underline{K_2Cr_2O_7}$ (g) $\underline{NH_4^+}$

(h) $\underline{HClO_4}$ (i) $\underline{HCl}$

→ (a) $\underline{CH_4}$

$x + 4(+1) = 0$

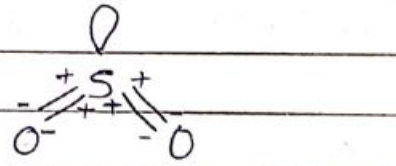
$x = -4$



(b) $\underline{SO_2}$

$x + 2(-2) = 0$

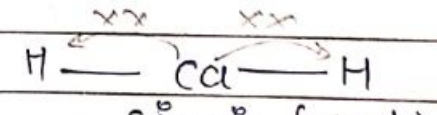
$x = +4$



(c) $\underline{CaH_2}$

$x + 2(+2) = 0$

$x = -2$



(ionic bond)

(d) $\underline{OF_2}$

$x + 2(-1) = 0$

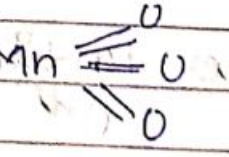
$x = +2$

(e) $\underline{KMnO_4}$

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

$1 + x - 8 = 0$

$x = +7$

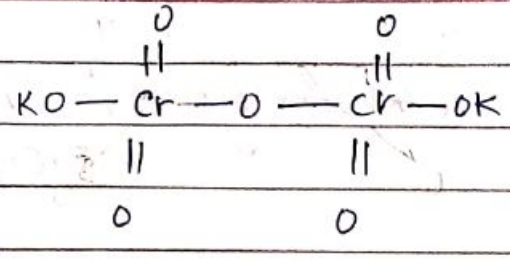


(f) $K_2Cr_2O_7$ (orange) (K dichromate)

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

$2+2x-14=0$

$2x=+12$. $x=+6$.

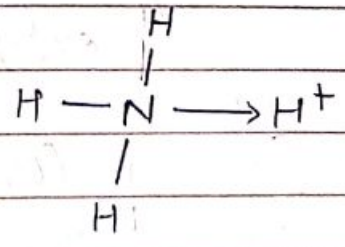


(g) NH_4^+

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

$x+4=1$

$x=-3$

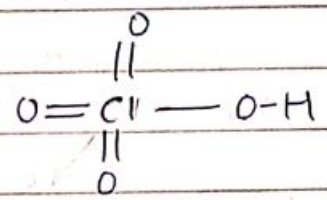


(h) $HClO_4$

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

$1+x-8=0$

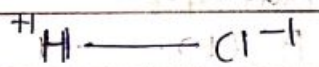
$x=+7$



(i) HCl

$1(+1)+x=0$

$x=-1$



Q:2 Find oxidation number of (a) chlorine in flowing:

(I) HCl (II) Cl_2 (III) $HClO_2$ (IV) $HClO_3$ (V) $HClO_4$

→ (I) $H-Cl=-1$ (II) $Cl-Cl=0$.

(III) $+1+x-4=0$ (IV) $1+x-6=0$

$x=+3$ $x=+5$

(V) $1+x-8=0$

$x=+7$.

NOTE: I) p-block elements show variable oxidation number.

II) maximum O.N for any group can be calculated by:

Group no. − 10. eg: Halogen 17−10 = +7

* halogen family show ON till +7 except fluorine (-1).

* chalcogen family show ON till +6 except oxygen.

* Osmium (Os) and Xenon (Xe) can show maximum ON of +8.

$OsO_4 \longrightarrow x - 8 = 0$

$x = +8$

Q:3 Find oxidation for the following underlined elements:

(a) $H_2\underline{S}O_5$ (caroacid) (b) $H\underline{N}O_4$ (per nitric acid)

(c) $Na_2\underline{S_4}O_6$ sodium tetrathionate

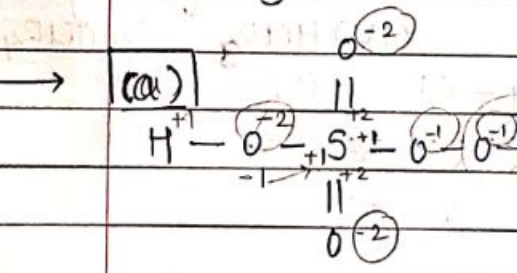
(d) $\underline{Br_3}O_8$

(e) $\underline{Cr}O_5$ (chromic oxide) (f) $CH_3\underline{COOH}$ (acetic acid)

(g) $\underline{Fe_3}O_4$ (h) $\underline{Pb_3}O_4$

(i) $\underline{K}I_3$ (j) $\underline{Mg_2}P_2O_7$

→ (a)



$2 + x - 10 = 0$

$x = +8$ X

O.N of S = +6

O.N of O = -1 and -2

avg. O.N of 'O' = $\frac{-8}{5} = -1.6$

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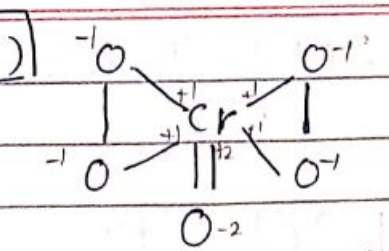
(e) 

$x - 10 = 0$

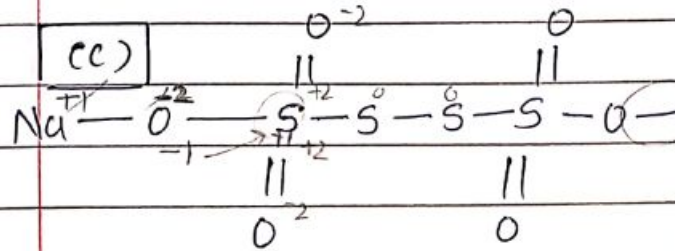
$x = +10$ (X)

(butterfly struc.)

O.N of 'Cr' = +6

O.N of 'O' = +1 & -2

avg. ON of 'O' $= \frac{-6}{5} = -1.2$

(c) 

not A bond (its ionic)

$2 + 4x - 12 = 0$

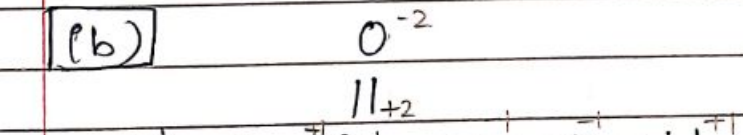
$4x = 10$

$x = +2.5$ (X)

O.N of 'S' = +5

O.N of 'O' = -1 & -2

avg. ON of 'O' =

(b) 

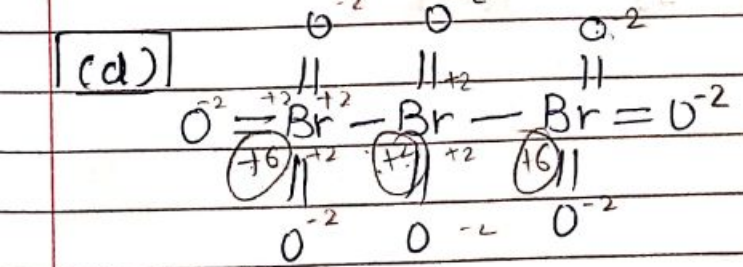
$1 + x - 8 = 0$

$x = +7$ (X)

O.N of 'N' = +5

O.N of 'O' = -1 & -2

avg. ON of 'O' $= \frac{-5}{4} = -1.25$

(d) 

$3x - 16 = 0$

$3x = 16$

$x = +5.3$ (X)

O.N of 'Br' = +6, +4

O.N of 'O' = -2

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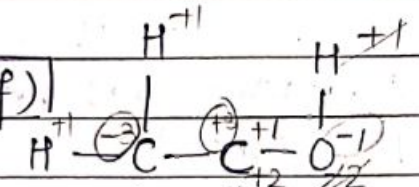
(f) $C_2H_4O_2$



$2x + 4 - 4 = 0$

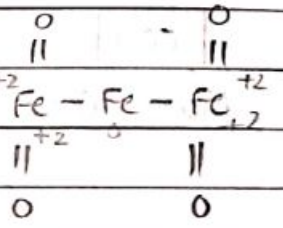
$x = 0$

ONO rule: 'C' = +3 ; −3 (avg.) = 0

mix oxide

(g) $Fe_3O_4 \longrightarrow FeO + Fe_2O_3$

+2 +3



$3x - 8 = 0$

$x = 8/3 \rightarrow +2.6$

$[x - 2 = 0 \;\; x = +2]$ $[2x - 6 = 0 \;\; x = +3]$

mix oxide

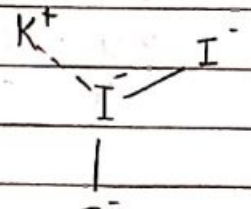
(h) $Pb_3O_4 \longrightarrow PbO + Pb_2O_3$

$[x - 2 = 0 \;\; x = +2]$ $[2x - 6 = 0 \;\; x = +3]$

$3x - 8 = 0$

$x = \frac{8}{3} \rightarrow +2.6$

(i) $KI_3 \longrightarrow K^+ \; I_3^-$ $\quad K = +1$



$(+1) + 3x = 0$

$3x = -1 \quad x = -1/3$ (I).

$\longrightarrow KI \cdot I_2$

$I = \frac{-1 + 0 + 0}{3} \rightarrow -1/3$

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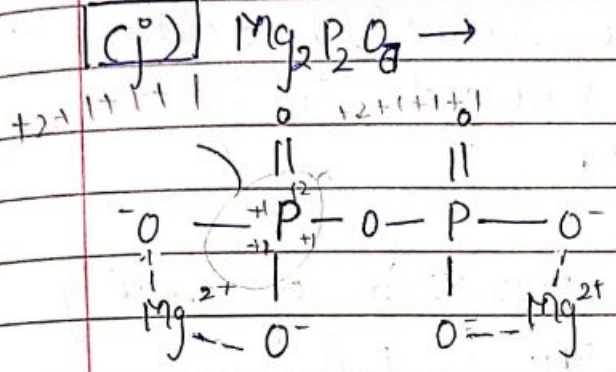
FeO = ferrous oxide Fe_2O_3 = ferric oxide

PbO = lead (II) oxide Pb_2O_3 = lead (III) oxide

c) $Mg_2P_2O_6 \rightarrow$



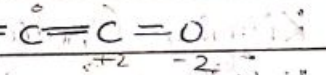
$2x + 4 - 14 = 0$

$2x = 10$

$x = +5$

ON = +5

C_3O_2 Ccarbon suboxide



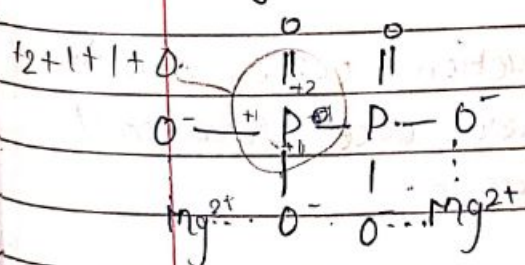
$\rightarrow 3x - 4 = 0$

$x = 4/3$

$+2 + 2 + 0 = 4$

$3 \quad\quad 3$

$Mg_2P_2O_6 \rightarrow$



$2x + 4 - 12 = 0$

$2x = 8$

$x = +4$

ON = +4

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Oxidising agent: (Oxidant)

- those reagent in redox reaction oxidises other reagent and it gets reduced (itself) are oxidising agent or oxidant.
- during redox reaction, reduction of oxidising agent is carried out. (electron hungry)
- more the oxidation state, oxidising property of reagent arises.

eg: $KMnO_4$ (potassium permanganate);
$K_2Cr_2O_7$ (potassium dichromate);
H_2SO_4 (sulphuric acid);
HNO_3 (nitric acid)

} strong oxi. agent

Reducing agent: (Reductant)

- those reagent in redox reaction which oxidises in itself and reduces other are reductant or reducing agent or antioxidant.
- during redox reaction, oxidation of reducing agent is carried out.

eg: $LiAlH_4$ (lithium aluminium hydride);
$NaBH_4$ (sodium boron hydride).

} strong red. agent.

Oxidation	Reduction
- reducing agent	- oxidising agent
- reductant	- oxidant
- antioxidant	- c
- oxidised	- reduced.

Q:4 $CH_4 + 2O_2 \longrightarrow \overset{-2}{CO_2} + 2H_2O$

⇒ ON: $CH_4 \rightarrow C = -4 \quad O_2 = 0 \quad H = +1$

$CO_2 \rightarrow C = +4 \; ; \; -2(O_2) \quad H = +1 \quad O = -1$

Oxidation = C (gains O). R.ag = C

Reduction = O_2 O.ag = O

reduction (O.agent) (reduced)

$CH_4 + O_2 \longrightarrow CO_2 + H_2O$

oxidation (R.agent) (oxidised)

Types of redox reaction:

1.) Combination redox reaction:

reduction

$\rightarrow H_2 + Cl_2 \longrightarrow 2HCl$

oxidation.

reduction

$\rightarrow PCl_3 + Cl_2 \longrightarrow PCl_5$

oxidation.

Ox = e⁻ loss
red = e⁻ gain

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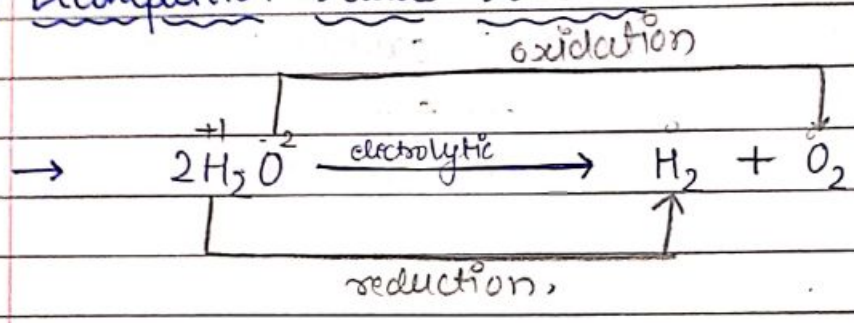
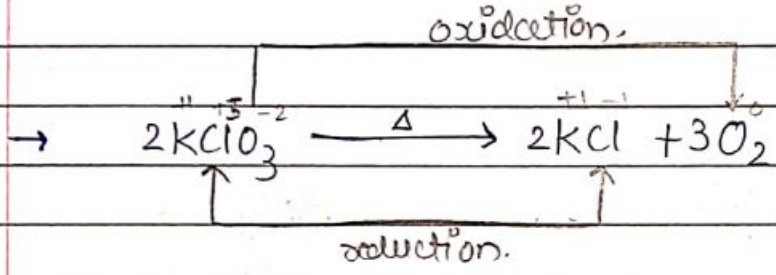
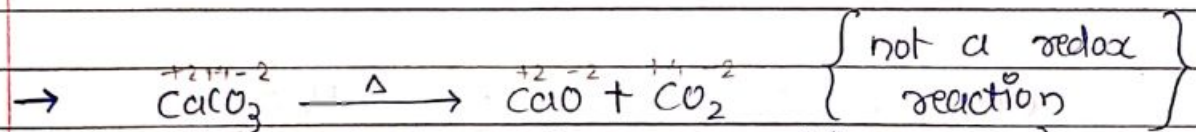
II.) Decomposition redox reaction:

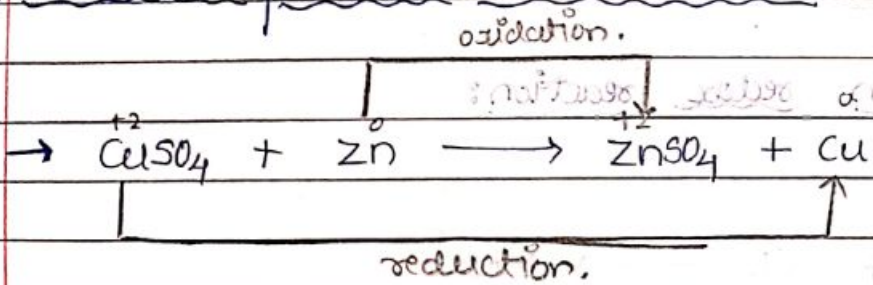


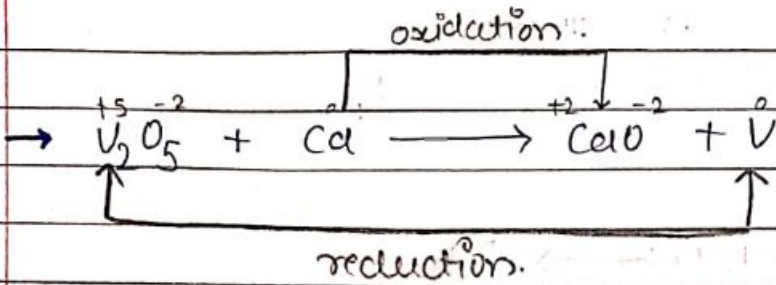




(just a thermal decomposition rxn).

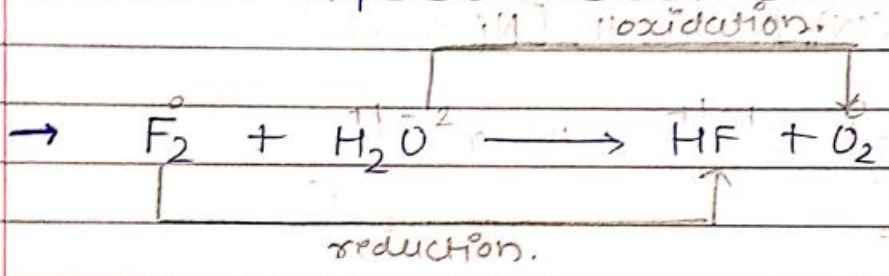
III.) Metal displacement redox reaction:





NOTE:	$(SO_4) \longrightarrow -2$		$(H_2O) \longrightarrow 0$	
	$(NH_4) \longrightarrow +1$		$(NH_3) \longrightarrow 0$	
nitric	HNO_3	$(NO_3) \longrightarrow -1$	carbonyl	$(CO) \longrightarrow 0$
nitrous	HNO_2	$(NO_2) \longrightarrow -1$		(C_2O_4)
	HCN	$(CN) \longrightarrow -1$		or $\longrightarrow -2$
		$(PO_4) \longrightarrow -3$		COO^-
		$(NO) \longrightarrow +1$		COO^-

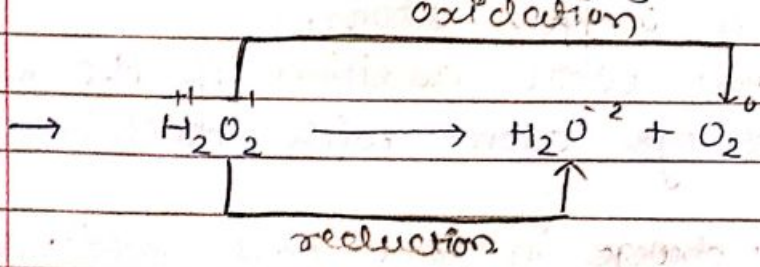
IV.) Non-metal displacement redox reaction:

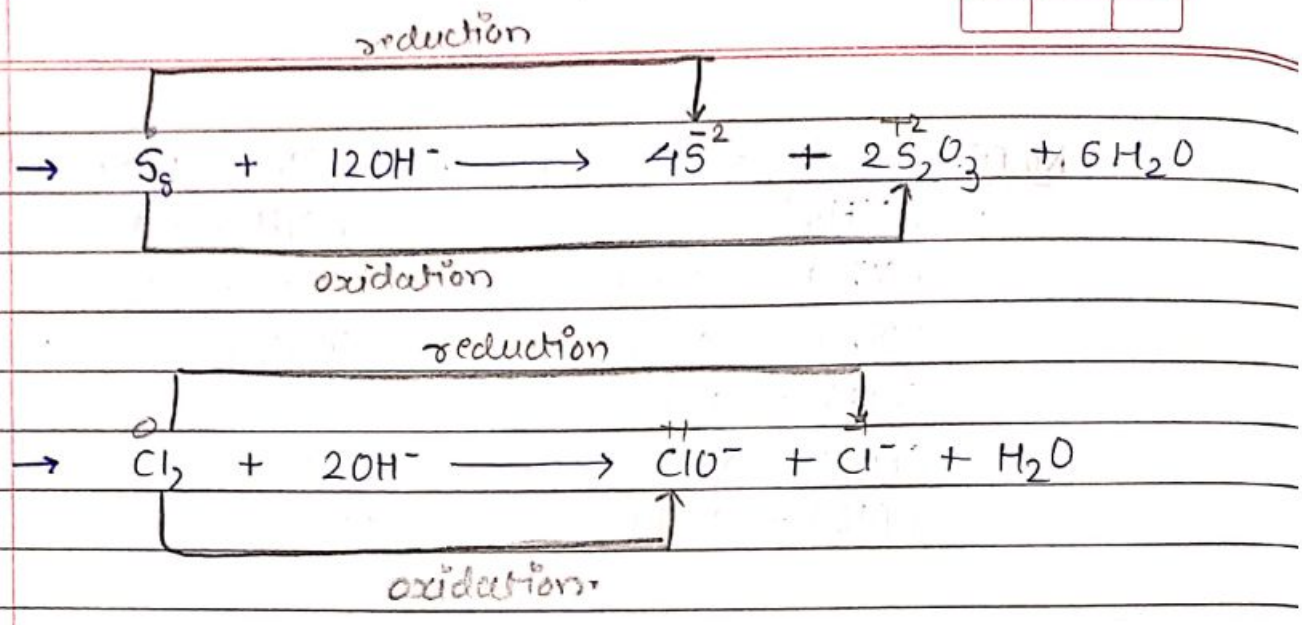


V.) Disproportionation redox reaction:

- those redox reaction in which an element in one oxidation state is simultaneously oxidized and reduced.

- one of the reactant in a disproportion reaction always contain element existing in 3 oxidation state.





Q:5 Find Oxidation no.:

I) $K_4[Fe(CN)_6]$ II) $[Ni(CO)_4]^{+2}$

→ I) $+4 + x - 6 = 0$ II) $x + 0 = +2$

$x = +2$ $x = +2$

Balancing of redox reaction:

★NCERT (1.) Ion-electron method (half ran method):

Rule:-
1. convert redox reaction to ionic form.
2. divide redox reaction in oxidation half reaction and reduction half reaction.
3. balance the metal / non-metal except hydrogen and oxygen atom.
4. balance oxygen with addition of H_2O molecule.
5. balance hydrogen atom with addition of H^+ ion.
6. balance the charge in both half reactions by adding electron (e^-).
7. after balancing the charge, now add both

half reaction, to remember that e^- can be cancelled out.

- this method can be explained by following eg's: (in acidic medium)

eg: 1) $Fe^{2+}_{(aq)} + Cr_2O_7^{2-}_{(aq)} \xrightarrow{H^+} Fe^{+3}_{(aq)} + Cr^{+3}_{(aq)}$

(+6 over $Cr_2O_7^{2-}$; oxidation: $Fe^{2+} \to Fe^{+3}$; reduction: $Cr_2O_7^{2-} \to Cr^{+3}$)

- $Fe^{+2} \longrightarrow Fe^{+3}$ (oxi. 1/2 rxn)

eq(I) $\Longrightarrow$ $Fe^{+2} \longrightarrow Fe^{+3} + e^-$ (balance charge)

- $Cr_2O_7^{2-} \longrightarrow Cr^{+3}$ (red. 1/2 rxn).

$\Rightarrow Cr_2O_7^{2-} \longrightarrow 2Cr^{+3}$ (balance metal).

$\Rightarrow Cr_2O_7^{2-} \longrightarrow 2Cr^{+3} + 7H_2O$ (balance oxygen).

$\Rightarrow Cr_2O_7^{2-} + 14H^+ \longrightarrow 2Cr^{+3} + 7H_2O$ (balance hydrogen)

$-2 + 14 = +12$; $2Cr^{+3}$: $+6$.

eq(II) $\Longrightarrow Cr_2O_7^{2-} + 6e^- + 14H^+ \longrightarrow 2Cr^{+3} + 7H_2O$ (bal. e^-)

- to cancel out electrons 6 is to be multiplied to reaction (I).

$6Fe^{+2} \longrightarrow 6Fe^{+3} + 6e^-$ — eq(III)

(adding eq (II) + eq(III))

$\Longrightarrow Cr_2O_7^{2-} + 14H^+ + 6Fe^{+2} \longrightarrow 2Cr^{+3} + 7H_2O + 6Fe^{+3}$

(it is acidic due to H^+ ion presence).

⇝ In Acidic medium:

Q:6) Balance following in acidic medium:

a) $MnO_4^- + SO_2 \longrightarrow Mn^{2+} + HSO_4^-$

reduction

oxidation

$\Rightarrow$ $MnO_4^- \longrightarrow Mn^{2+}$ (mass balanced) (red. ½ rxn)..

- $MnO_4^- \longrightarrow Mn^{2+} + 4H_2O$ (O. balance)
- $MnO_4^- + 8H^+ \longrightarrow Mn^{2+} + 4H_2O$ (H balance)

(+7) (+2).

$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$ (charge balance) (eq(I))

$\Rightarrow$ $SO_2 \longrightarrow HSO_4^-$ (mass balance) (oxi. ½ rxn)

- $SO_2 + 2H_2O \longrightarrow HSO_4^-$ (balance O)
- $SO_2 + 2H_2O \longrightarrow HSO_4^- + 3H^+$ (balance H)

(0) +(0)) (-1 + 3 = +2)

$SO_2 + 2H_2O \longrightarrow HSO_4^- + 3H^+ + 2e^-$ (charge bal.) (eq(II))

$\Rightarrow$ eq(I)×2 eq(II)×5

$2MnO_4^- + 16H^+ + 10e^- \longrightarrow 2Mn^{2+} + 8H_2O$

$5SO_2 + 10H_2O \longrightarrow 5HSO_4^- + 15H^+ + 10e^-$

$\Rightarrow$ $2MnO_4^- + H^+ + 5SO_2 + 2H_2O \longrightarrow 2Mn^{2+} + 5HSO_4^+$

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(b) $FeSO_4 + KMnO_4 + H_2SO_4 \longrightarrow Fe_2(SO_4)_3 + MnSO_4 + H_2O + K_2SO_4$

$\Longrightarrow$ $\overset{+2}{Fe}SO_4 + \overset{+1}{K}\overset{+7}{Mn}O_4 + H_2SO_4 \longrightarrow \overset{+3}{Fe_2}(SO_4)_3 + \overset{+2}{Mn}SO_4 + H_2O + K_2SO_4$

oxidation

reduction.

$\Longrightarrow$ $Fe^{2+} + MnO_4^- \longrightarrow 2Fe^{+3} + Mn^{+2}$ (ionic).

red.

oxi.

- $Fe^{2+} \longrightarrow Fe^{3+}$ (oxi. 1/2 rxn).

  $Fe^{2+} \longrightarrow Fe^{3+} + e^-$ (charge) — eq(1)

- $MnO_4^- \longrightarrow Mn^{+2}$ (red. 1/2 rxn).

  $MnO_4^- \longrightarrow Mn^{+2} + 4H_2O$ (O bal)

  $MnO_4^- + 8H^+ \longrightarrow Mn^{+2} + 4H_2O$ (H bal).

  (+7) (+2)

eq(11) $MnO_4^- + 5e^- + 8H^+ \longrightarrow Mn^{+2} + 4H_2O$ (charge)

∴ eq(1) x 5. eq(11) x 1

- $5Fe^{2+} \longrightarrow 5Fe^{3+} + 5e^-$
- $MnO_4^- + 5e^- + 8H^+ \longrightarrow Mn^{+2} + 4H_2O$

$\Longrightarrow 5FeSO_4 + KMnO_4 + 4H_2SO_4 \longrightarrow \frac{5}{2}Fe_2(SO_4)_3 + MnSO_4 + \frac{1}{2}K_2SO_4 + 4H_2O$

Or x2 $10FeSO_4 + 2KMnO_4 + 8H_2SO_4 \longrightarrow 5Fe_2(SO_4)_3 + 2MnSO_4 + K_2SO_4 + 8H_2O$

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→ In basic medium:

- can be understood by following redox reaction:

Q: 7) (a) $MnO_4^- + I^- \longrightarrow MnO_2 + I_2$

(+7, -1, +4, 0; oxidation, reduction)

→ $2I^- \longrightarrow I_2$ (balance m) (oxi. ½ rxn)

eq(1) $2I^- \longrightarrow I_2 + 2e^-$ (bala. charge).

→ $MnO_4^- \longrightarrow MnO_2$ (red. ½ rxn) (m bal.)

$4H^+ + MnO_4^- \longrightarrow MnO_2 + 2H_2O$ (O bal)

(+7) → (+3), (+3) 0

$4H^+ + MnO_4^- + 3e^- \longrightarrow MnO_2 + 2H_2O$ (e⁻ bal) H⁺

adding $4OH^-$ to bala. with $4H^+$

basic

$(4H^+ + 4OH^-) + MnO_4^- + 3e^- \longrightarrow MnO_2 + 2H_2O + 4OH^-$ — eq2

$4H_2O + MnO_4^- + 3e^- \longrightarrow MnO_2 + 2H_2O + 4OH^-$

$2H_2O + MnO_4^- + 3e^- \longrightarrow MnO_2 + 4OH^-$ — eq 3

eq(1) × 3 (bal +3e⁻) & eq(III) × 2 (bal +2e⁻)

$6I^- \longrightarrow 3I_2 + 6e^-$

$4H_2O + 2MnO_4^- + 6e^- \longrightarrow 2MnO_2 + 8OH^-$

$6I^- + 2MnO_4^- + 4H_2O \longrightarrow 3I_2 + 2MnO_2 + 8OH^-$

net charge: $-6 - 2 = -8$ (−8)

(b) $ClO^- + CrO_2^- + OH^- \longrightarrow Cl^- + CrO_4^{2-} + H_2O$

Q:8 Balance follow in basic medium by ½ reaction method.

→ $P_4 \xrightarrow{OH^-} PH_3 + H_2PO_2^-$

red. ox.

$\Rightarrow P_4 \longrightarrow 4H_2PO_2^-$ (1).

$P_4 + 8H_2O \longrightarrow 4H_2PO_2^-$ (O).

$P_4 + 8H_2O \longrightarrow 4H_2PO_2^- + 8H^+$ (H).

adding OH^-

$P_4 + 8H_2O + 8OH^- \longrightarrow 4H_2PO_2^- + 8H^+ + 8OH^-$

(-3) $P_4 + 8H_2O + 8OH^- \longrightarrow 4H_2PO_2^- + 8H_2O$ (-1)

$\therefore P_4 + 8OH^- \longrightarrow 4H_2PO_2^- + 4e^-$ (I-) eq(I)

$\Rightarrow P_4 \longrightarrow 4PH_3$ (1)

$P_4 + 12H^+ \longrightarrow 4PH_3$ (H). add OH^-

$P_4 + 12H^+ + 12OH^- \longrightarrow 4PH_3 + 12OH^-$

$P_4 + 12H_2O \longrightarrow 4PH_3 + 12OH^-$ (-12)

$\therefore P_4 + 12H_2O + 12e^- \longrightarrow 4PH_3 + 12OH^-$ (I-) eq(II)

$\therefore$ eq(I) × 3 eq(II) × 1.

$3P_4 + 24OH^- \longrightarrow 12H_2PO_2^- + 12e^-$

$P_4 + 12H_2O + 12e^- \longrightarrow 4PH_3 + 12OH^-$

$\therefore 4P_4 + 12OH^- + 12H_2O \longrightarrow 12H_2PO_2^- + 4PH_3$