ULTIMATE KCET

CRASH COURSE 2026

Chemistry | Lecture - 01

1-3

Redox reaction

By – Sreeja Ma'am

Recap *of previous lecture*

1. Coordination compound - 2

Topics *to be covered*

Redox reaction


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Redox reactions

- Types of reactions, Redox titration of $KMnO_4$/ FAS – 2025
- Deciding reducing agent in a chemical equation – 2023,2013
- Oxidation state calculation – 2023,2022,2020,2019,2011,2010,2008
- 2007,2006
- Definition of oxidation – 2021,2017
- Calculating coefficient in ionic equation – 2018,

Redox reactions → Electron concept

Oxidation

- → Addition of oxygen
- → Addition of Electronegative element
- → Remove Hydrogen
- → Remove electropositive element

Reduction

- → Removal of oxygen
- → Removal of Electronegative element
- → Addition of +1
- → Addition of Electropositive element

Nonmetals ← electronegative element → N^{3-}, O^{2-}, $F^{\ominus}$, $Cl^{\ominus}$, $Br^{\ominus}$, $I^{\ominus}$

Metals ← Electropositive element → Li^{+}, Na^{+}, K^{+}, Rb^{+}, Fe^{+2}, Hg^{+2}

$(S \rightarrow S^{-2})$ Reduction. $S^{-2} \rightarrow S^{0}$ → oxidation.

Redox reaction → Simultaneous oxidation-Reduction reaction → Redox reaction

→ Oxidation State number.

No. of electron lost or gained during the chemical reaction

O^{-2}, S^{-2}, N^{-3} (gain of e⁻)

Na^{+}, Hg^{2+}, Al^{+3} (loss of e⁻)

Oxidation

→ loss of electron

→ Ex: $\overset{0}{C} \longrightarrow \overset{+4}{C}$ (lost 4e⁻)

$S^{-2} \longrightarrow \overset{0}{S}$ (lost 2e⁻)

→ Oxidation no. increases

Reduction

gain of electron

$\overset{0}{S} \longrightarrow S^{-2}$ (gained 2e⁻)

$\overset{+7}{Mn}O_4^- \xrightarrow[\text{gain}]{5e^-} \overset{+2}{Mn}$ (gained 5e⁻)

Oxidation no decreases

Calculation on Oxidation number.

$$\overset{0}{Mg} + \overset{0}{O_2} \longrightarrow \overset{+2\,-2}{MgO}$$

gain of e^- → reduction

0. NO ↑ oxidation — loss of e^-

BaO	BaO_2
$\overset{+2}{Ba}\ \overset{-2}{O}$	~~+2~~ ~~-2~~ Ba O_2

#1. $\overset{0}{Mg}$ $\overset{0}{Cu}$ $\overset{0}{Zn}$ $\overset{0}{H_2}$ $\overset{0}{Fe}$ $\overset{0}{N_2}$

→ Elemental or Homoatomic molecule has zero oxidation state

#2. → Charge on single ion is its oxidation state.

ex: Al^{+3}, Zn^{+2}, Mg^{2+}

Mn^{+2}

Rule 4

Group 1 elements

Li, Na, K, Rb, Cs → +1

Group 2 elements → +2

Be, Mg, Ca, Sr, Ba → (+2)

Group 13 →

Al^{+3}, Ga^{+3}, In^{+3}, Tl^{+3}

Rule → 3

* oxide → O^{-2} ✓
* peroxide → O^{-1}
* superoxide → $O^{-1/2}$.

* Superoxides

KO_2, RbO_2 } $\overset{+1}{K}\ \overset{-1}{O_2}$ { $\overset{-1/2}{O}\ \overset{-1/2}{O}$ }

Rule 5

Only flourine shows −ve oxidation in all the compounds.

Exceptional case for oxygen (V.Imp)

$\overset{+1}{O_2}F_2 \rightarrow 2x + (-2) = 0$

$2x = 2,\ x = +1$

$\overset{+2}{O}F_2$

$x + (-1 \times 2) = 0 \rightarrow x = +2$

Rule #6

$F^{\ominus} \longrightarrow -1$

$Cl^{\ominus} \longrightarrow -1, +1, +3, +5, +7$

$Br^{\ominus} \longrightarrow -1, +1, +3, +5, +7$

$I^{\ominus} \longrightarrow -1, +1, +3, +5, +7$

$BrO_3^{\ominus} \rightarrow$ Bromate

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

$x - 6 = -1$

$x = -1 + 6 = +5$

$IO_3^{\ominus}$ → iodate

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

$x - 6 = -1$

$x = +5$

Cl_2O_7

$2x + (-2 \times 7) = 0$

$2x - 14 = 0$

$x = 14/2 = +7$

Rule #7:

→ Hydrogen + nonmetal ⟶ +1

→ Hydrogen + metal ⟶ −1

H

Nonmetals → +1

NH_3

CH_4

PH_3

Metals → −1

ZnH_2^{-1}

CaH_2^{-1}

NaH^{-1}

Inert pair effect

$Sn^{+2} \longrightarrow Sn^{4+} + 2e^{-}$ Sn^{+4}, Sn^{+2}

Inert pair effect → Chemically unreactive.

$Pb^{+4} \longrightarrow$ Oxidizing agent

$Pb^{4+} + 2e^{-} \longrightarrow Pb^{+2}$

e⁻ accept, ready to → O.A

13		14		15
2nd: B	$2s^2 2p^1$	C	$2s^2 2p^2$	N
3rd: Al	$3s^2 3p^1$	Si	$3s^2 3p^2$	P
4th: Ga	$3d^{10} 4s^2 4p^1$	Ge	$3d^{10} 4s^2 4p^2$	As
5th: In	$4d^{10} 5s^2 5p^1$	Sn	$4d^{10} 5s^2 5p^2$	Sb → +4, +2
6th: Tl	$4f^{14} 5d^{10} 6s^2 6p^1$	Pb	$4f^{14}, 5d^{10}, 6s^2 6p^2$	Bi

Tl → +1   Pb → +2

Group IIIA (13) ($ns^2 np^1$)	O.S
B	+3
Al	+3
Ga	+3, +1
In	+3, +1
Tl	+3, +1

Group IVA (14) ($ns^2 np^2$)	O.S
C	+4
Si	+4
Ge	+4, +2
Sn	+4, +2
Pb	+4, +2

Group VA (15) ($ns^2 np^3$)	O.S	
2nd N	+3	→ $2s^2 2p^3$
3rd P	+3, +5	→ $3s^2 3p^3$
4th As	+3, +5	→ $3d^{10} 4s^2 4p^3$
5th Sb	+3, +5	→ $4d^{10} 5s^2 5p^3$
6th Bi	+3, +5	→ $4f^{14} 5d^{10} 6s^2 6p^3$

Cr (+6), Mo (+6), W (+6) → R.A → Ready to lose e⁻

Oxidation/reduction – Oxidation number

Oxidation: An increase in the oxidation number of the element in the given substance.

Reduction : A decrease in the oxidation number of the element in the given substance.

Oxidising agent: A reagent which can increase the oxidation number of an element in a given substance. These reagents are called as oxidants also.

Reducing agent: A reagent which lowers the oxidation number of an element in a given substance. These reagents are also called as reductants.

Redox reactions: Reactions which involve change in oxidation number of the interacting species

Justify that the reaction:

$2Cu_2O_{(s)} + Cu_2S_{(s)} \rightarrow 6Cu_{(s)} + SO_{2(g)}$ is a redox reaction.

Identify the species oxidised/reduced, which acts as an oxidant and which acts as a reductant.

$$2Cu_2O_{(s)} + Cu_2S_{(s)} \rightarrow 6Cu_{(s)} + SO_{2(g)}$$

We therefore, conclude that in this reaction copper is reduced from +1 state to zero oxidation state and sulphur is oxidised from –2 state to +4 state.

The above reaction is thus a redox reaction.

Stock notation

Ferrous → Fe^{+2} Mercurous → Hg^{+1}

Ferric → Fe^{+3} Mercuric → Hg^{+2}.

Cuprous → Cu^{+1}

Cupric → Cu^{+2}.

Stock notation: According to this, the oxidation number is expressed by putting a Roman numeral representing the oxidation number in parenthesis after the symbol of the metal in the molecular formula.

- Aurous chloride as Au(I)Cl. → Au^{+1}
- Auric chloride as Au(III)Cl_3. → Au^{+3}
- Stannous chloride Sn(II)Cl_2 → Sn^{+2}
- Stannic chloride Sn(IV)Cl_4. → Sn^{+3}

$AuCl_3$ → Au(III)Cl_3

$x + (-1 \times 3) = x - 3 = 0$

$x = +3$

Au^{+3}

Cuprous chloride

↓

Cu(I)Cl

Cupric chloride

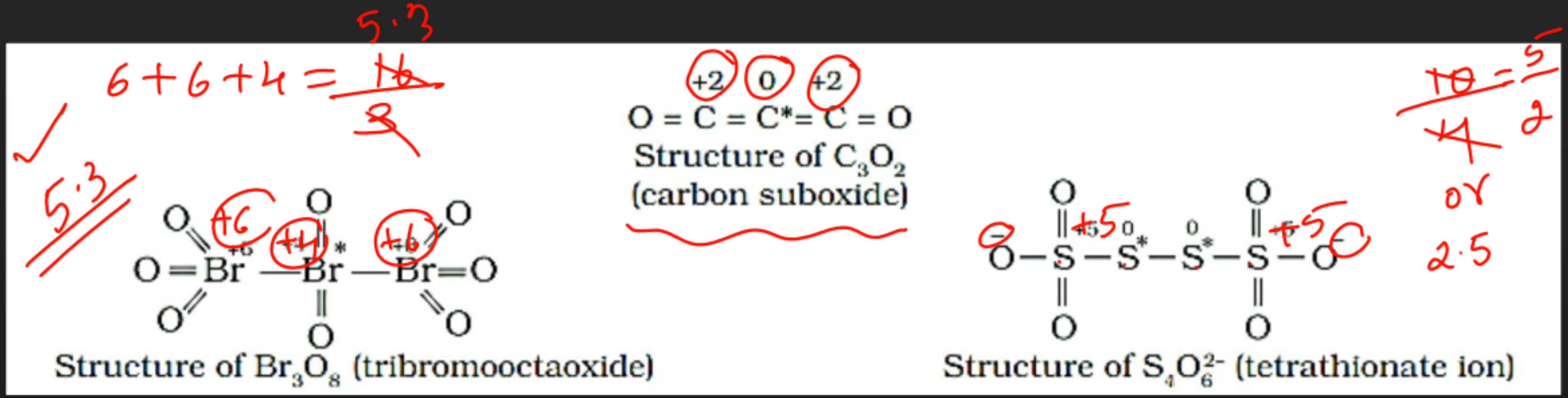
Cu(II)Cl_2

Cu^{+1} Cl^{-1}

$\overset{+2}{Cu}$ $\overset{-1}{Cl}$

$CuCl_2$

Fractional Oxidation Number

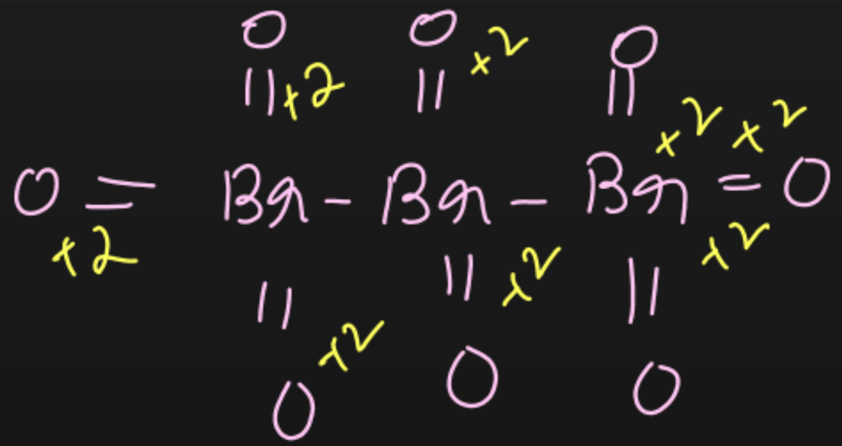


Structure of Br_3O_8 (tribromooctaoxide)

Structure of C_3O_2 (carbon suboxide)

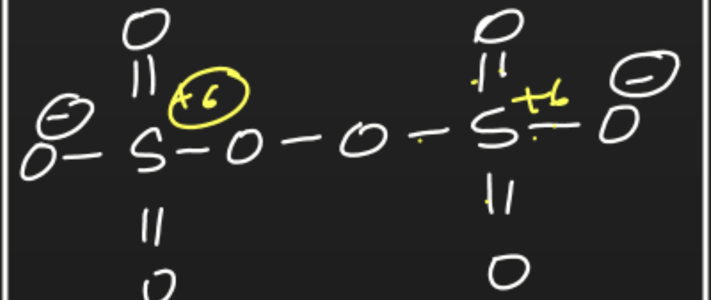
Structure of $S_4O_6^{2-}$ (tetrathionate ion)

Br_3O_8



$S_2O_8^{2-}$

peroxodisulphate ion



$6+6 = \frac{12}{2} = 6$

Average oxidation State = +6

Types of Redox Reactions - Combination reactions

Combination reactions: More than two elements or compounds combining to form product is known as combination reaction.

$C_{(s)} + O_{2\,(g)} \longrightarrow CO_{2(g)}$ Redox

$3\overset{0}{Mg}_{(s)} + \overset{0}{N}_{2(g)} \longrightarrow Mg_3N_{2(s)}$

$CH_{4(g)} + 2\overset{0}{O}_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O\,(l)$

$CaO + H_2O \longrightarrow Ca(OH)_2$

Combination ✓ Redox ✗

Types of Redox Reactions - Decomposition reactions

Decomposition reactions: decomposition reaction leads to the breakdown of a compound into two or more components at least one of which must be in the elemental state.

$$2H_2O_{(l)} \longrightarrow 2H_{2(g)} + O_{2(g)}$$

$$2NaH_{(s)} \longrightarrow 2Na_{(s)} + H_{2(g)}$$

$$2KClO_{3(s)} \longrightarrow 2KCl_{(s)} + 3O_{2(g)}$$

Note: All decomposition reactions are not redox reactions. For example, decomposition of calcium carbonate is not a redox reaction.

$$CaCO_{3(s)} \longrightarrow CaO_{(s)} + CO_{2(g)}$$

Types of Redox Reactions - Displacement reactions

Displacement reactions:In a displacement reaction, an ion (or an atom) in a compound is replaced by an ion (or an atom) of another element.[ncert definition]

Displacement reactions fit into two categories:

- Metal displacement
- Non-metal displacement.

Metal displacement: A metal in a compound can be displaced by another metal in the uncombined state provided it occupies a position higher in the activity series.

$CuSO_{4(aq)} + Zn_{(s)} \rightarrow Cu_{(s)} + ZnSO_4$

$V_2O_{5\,(s)} + 5Ca_{(s)} \rightarrow 2V_{(s)} + 5CaO_{(s)}$

$TiCl_{4\,(l)} + 2Mg_{(s)} \rightarrow Ti_{(s)} + 2MgCl$

$Cr_2O_{3\,(s)} + 2Al_{(s)} \rightarrow Al_2O_{3\,(s)} + 2Cr_{(s)}$

how reactive High reactive

Application :In metallurgical operations for the isolation of certain metals from their respective ores.

Non-metal displacement: The nonmetal displacement redox reactions include hydrogen displacement.

Case-1:Displacement of hydrogen from water.

All alkali metals and some alkaline earth metals (Ca, Sr, and Ba) which are very good reductants, will displace hydrogen from cold water.

$$2Na_{(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}$$

$$Ca_{(s)} + 2H_2O_{(l)} \rightarrow Ca(OH)_2\ (aq) + H_{2(g)}$$

Less active metals such as magnesium and iron react with steam to produce dihydrogen gas.

$$Mg_{(s)} + 2H_2O_{(l,\ hot\ water)} \longrightarrow Mg(OH)_{2(s)} + H_{2(g)}$$

$$Fe_{(s)} + 2H_2O_{(g,\ steam)} \longrightarrow Mg(OH)_{2(s)} + H_{2(g)}$$

Case-2:Displacement of hydrogen from acids .

$$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2\ (aq)} + H_{2\ (g)}$$

- The power of these elements as oxidising agents decreases as we move down from fluorine to iodine in group 17 of the periodic table.
- Fluorine is so reactive that it attacks water and displaces the oxygen of water
- $2H_2O_{(l)} + 2F_{2(g)} \rightarrow 4HF_{(aq)} + O_{2(g)}$
- It is for this reason that the displacement reactions of chlorine, bromine and iodine using fluorine are not generally carried out aqueous solution.

$F^{\ominus}$, $Cl^{\ominus}$, $Br^{\ominus}$, $I^{\ominus}$ } oxidizing agent

$F^{\ominus} > Cl^{\ominus} > Br^{\ominus} > I^{\ominus}$

loss of $e^{\ominus}$ → oxidation

$\overset{-2}{H_2O} + \overset{0}{2F_2} \longrightarrow \overset{-1}{HF} + \overset{0}{O_2}$

Gain of $e^{\ominus}$

Disproportionation reactions : In a disproportionation reaction an element in one oxidation state is simultaneously oxidised and reduced.

The decomposition of hydrogen peroxide

$$2H_2O_{2\,(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)}$$

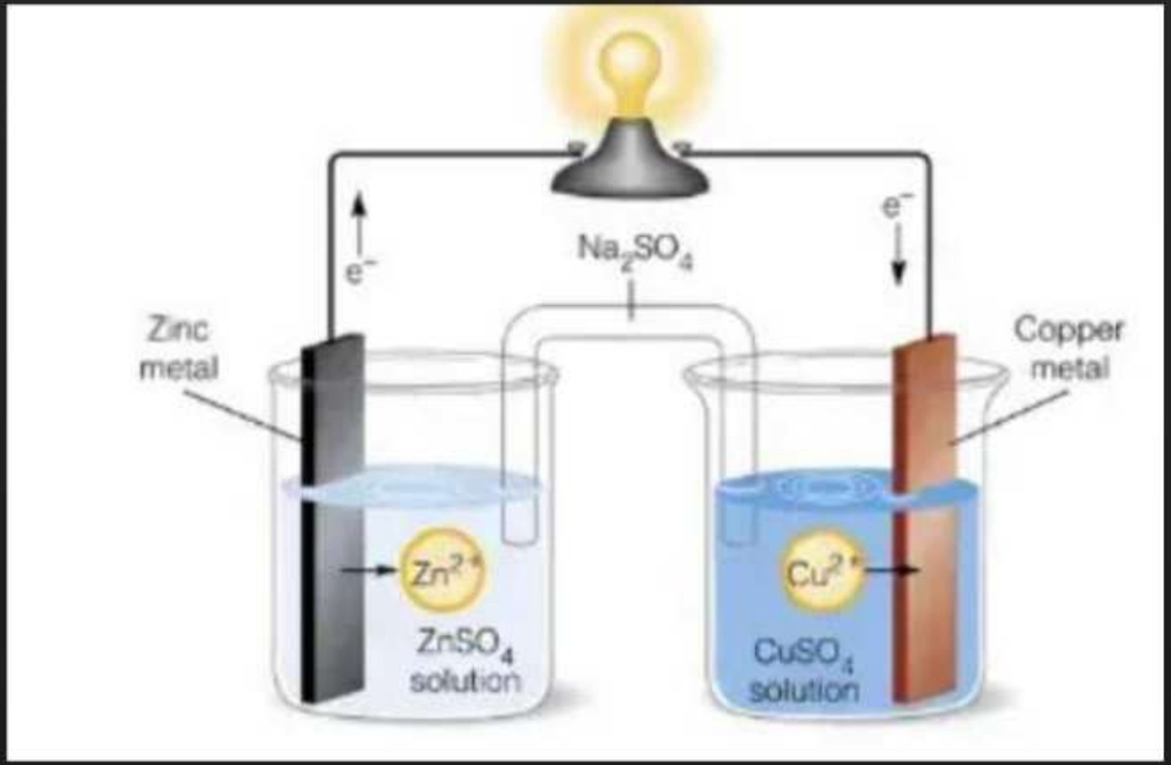
$$P_{4(s)} + 3OH^-_{(aq)} + 3H_2O_{(l)} \rightarrow PH_{3(g)} + 3H_2PO_2^-$$

$$S_{8(s)} + 12\,OH^-_{(aq)} \rightarrow 4S^{2-}_{(aq)} + 2S_2O_3^{\,2-}{}_{(aq)} + 6H_2O_{(l)}$$

$$Cl_{2\,(g)} + 2\,OH^-_{(aq)} \rightarrow ClO^-_{(aq)} + Cl^-_{(aq)} + H_2O_{(l)}$$

The hypochlorite ion (ClO^-) formed in the reaction oxidizes the colour-bearing stains of the substances to colorless compounds.

DANIEL CELL :



Na2SO4
e⁻
e⁻
Zinc metal
Copper metal
Zn2+
Cu2+
ZnSO4 solution
CuSO4 solution

Redox couple :

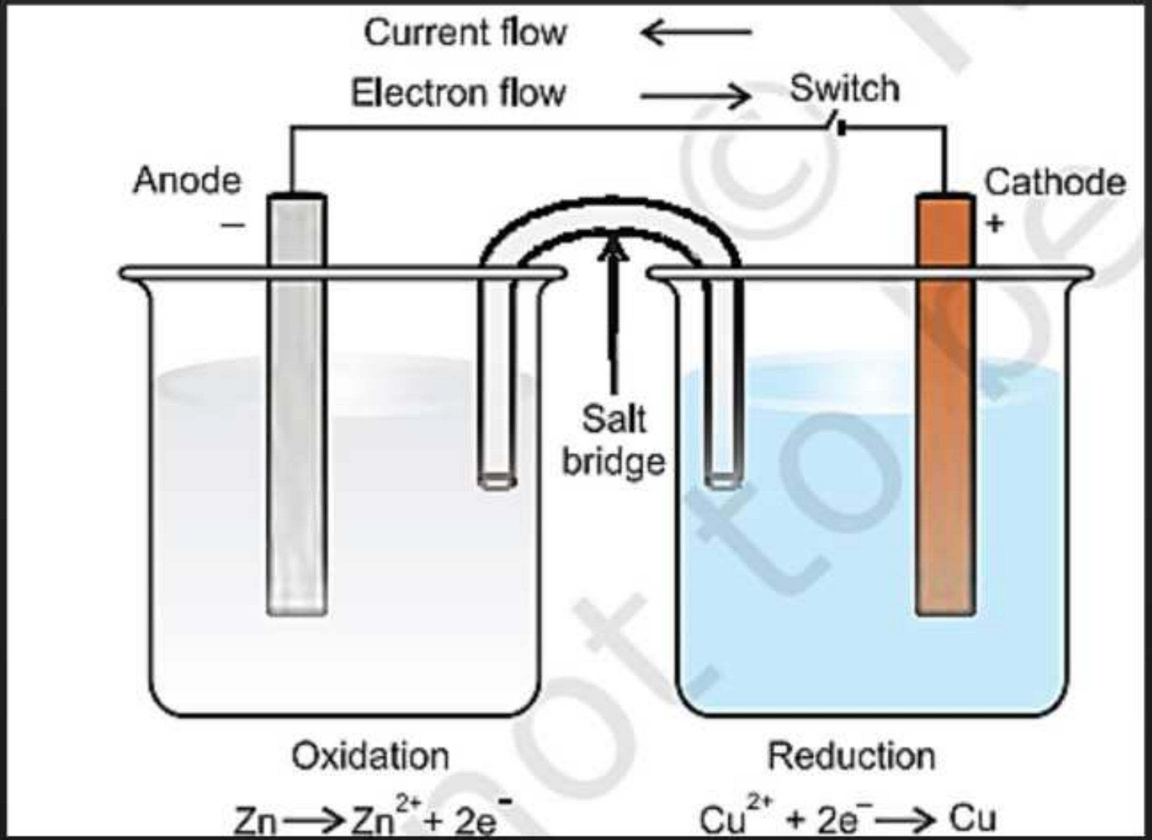
The oxidised and reduced forms of a substance taking part in an oxidation or reduction half reaction.

This is represented by separating the oxidised form from the reduced form by a vertical line or a slash representing an interface.

Zn^{2+}/Zn and Cu^{2+}/Cu.

$Zn/Zn^{2+} || Cu^{2+}/Cu$

Redox Reactions and Electrode Processes



Increasing strength of oxidising agent (↑)

Increasing strength of reducing agent (↓)

Reaction (Oxidised form + ne^-	→ Reduced form)	E° / V
$F_2(g) + 2e^-$	$\rightarrow 2F^-$	2.87
$Co^{3+} + e^-$	$\rightarrow Co^{2+}$	1.81
$H_2O_2 + 2H^+ + 2e^-$	$\rightarrow 2H_2O$	1.78
$MnO_4^- + 8H^+ + 5e^-$	$\rightarrow Mn^{2+} + 4H_2O$	1.51
$Au^{3+} + 3e^-$	$\rightarrow Au(s)$	1.40
$Cl_2(g) + 2e^-$	$\rightarrow 2Cl^-$	1.36
$Cr_2O_7^{2-} + 14H^+ + 6e^-$	$\rightarrow 2Cr^{3+} + 7H_2O$	1.33
$O_2(g) + 4H^+ + 4e^-$	$\rightarrow 2H_2O$	1.23
$MnO_2(s) + 4H^+ + 2e^-$	$\rightarrow Mn^{2+} + 2H_2O$	1.23
$Br_2 + 2e^-$	$\rightarrow 2Br^-$	1.09
$NO_3^- + 4H^+ + 3e^-$	$\rightarrow NO(g) + 2H_2O$	0.97
$2Hg^{2+} + 2e^-$	$\rightarrow Hg_2^{2+}$	0.92
$Ag^+ + e^-$	$\rightarrow Ag(s)$	0.80
$Fe^{3+} + e^-$	$\rightarrow Fe^{2+}$	0.77
$O_2(g) + 2H^+ + 2e^-$	$\rightarrow H_2O_2$	0.68
$I_2(s) + 2e^-$	$\rightarrow 2I^-$	0.54
$Cu^+ + e^-$	$\rightarrow Cu(s)$	0.52
$Cu^{2+} + 2e^-$	$\rightarrow Cu(s)$	0.34
$AgCl(s) + e^-$	$\rightarrow Ag(s) + Cl^-$	0.22
$AgBr(s) + e^-$	$\rightarrow Ag(s) + Br^-$	0.10
$2H^+ + 2e^-$	**$\rightarrow H_2(g)$**	**0.00**
$Pb^{2+} + 2e^-$	$\rightarrow Pb(s)$	−0.13
$Sn^{2+} + 2e^-$	$\rightarrow Sn(s)$	−0.14
$Ni^{2+} + 2e^-$	$\rightarrow Ni(s)$	−0.25
$Fe^{2+} + 2e^-$	$\rightarrow Fe(s)$	−0.44
$Cr^{3+} + 3e^-$	$\rightarrow Cr(s)$	−0.74
$Zn^{2+} + 2e^-$	$\rightarrow Zn(s)$	−0.76
$2H_2O + 2e^-$	$\rightarrow H_2(g) + 2OH^-$	−0.83
$Al^{3+} + 3e^-$	$\rightarrow Al(s)$	−1.66
$Mg^{2+} + 2e^-$	$\rightarrow Mg(s)$	−2.36
$Na^+ + e^-$	$\rightarrow Na(s)$	−2.71
$Ca^{2+} + 2e^-$	$\rightarrow Ca(s)$	−2.87
$K^+ + e^-$	$\rightarrow K(s)$	−2.93
$Li^+ + e^-$	$\rightarrow Li(s)$	−3.05

1. A negative E° means that the redox couple is a stronger reducing agent than the H^+/H_2 couple.
2. A positive E° means that the redox couple is a weaker reducing agent than the H^+/H_2 couple.

7.3.3 Redox Reactions as the Basis for Titrations

In acid-base systems we come across with a titration method for finding out the strength of one solution against the other using a pH sensitive indicator. Similarly, in redox systems, the titration method can be adopted to determine the strength of a reductant/oxidant using a redox sensitive indicator. The usage of indicators in redox titration is illustrated below:

(i) In one situation, the reagent itself is intensely coloured, e.g., permanganate ion, MnO_4^-. Here MnO_4^- acts as the self indicator. The visible end point in this case is achieved after the last of the reductant (Fe^{2+} or $C_2O_4^{2-}$) is oxidised and the first lasting tinge of pink colour appears at MnO_4^- concentration as low as 10^{-6} mol dm^{-3} (10^{-6} mol L^{-1}). This ensures a minimal 'overshoot' in colour beyond the equivalence point, the point where the reductant and the oxidant are equal in terms of their mole stoichiometry.

(ii) If there is no dramatic auto-colour change (as with MnO_4^- titration), there are indicators which are oxidised immediately after the last bit of the reactant is consumed, producing a dramatic colour change. The best example is afforded by $Cr_2O_7^{2-}$, which is not a self-indicator, but oxidises the indicator substance diphenylamine just after the equivalence point to produce an intense blue colour, thus signalling the end point.

(iii) There is yet another method which is interesting and quite common. Its use is restricted to those reagents which are able to oxidise I^- ions, say, for example, Cu(II):

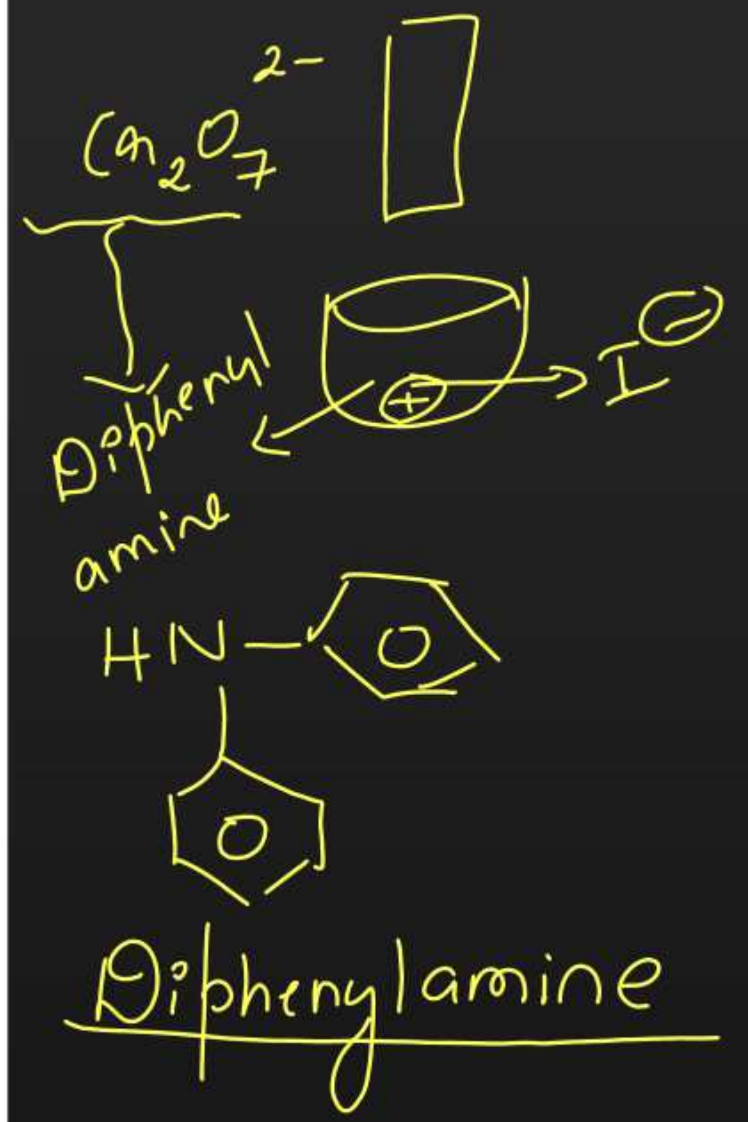
$$2Cu^{2+}(aq) + 4I^-(aq) \rightarrow Cu_2I_2(s) + I_2(aq) \quad (7.59)$$

This method relies on the facts that iodine itself gives an intense blue colour with starch and has a very specific reaction with thiosulphate ions ($S_2O_3^{2-}$), which too is a redox reaction:

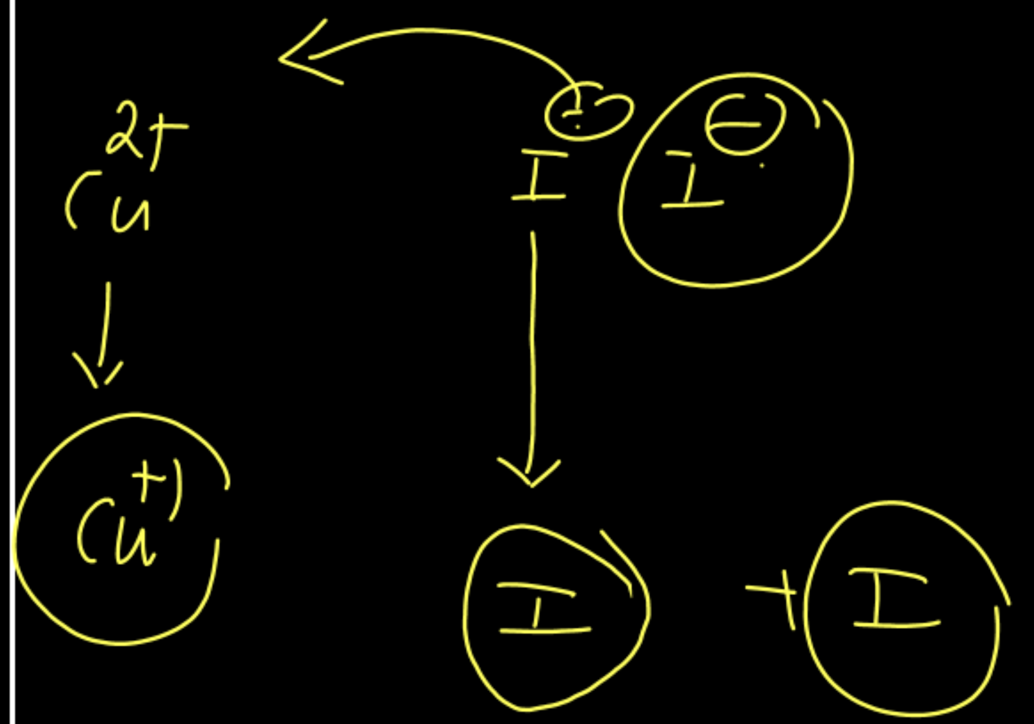
$$I_2(aq) + 2\,S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq) \quad (7.60)$$

I_2, though insoluble in water, remains in solution containing KI as KI_3.

On addition of starch after the liberation of iodine from the reaction of Cu^{2+} ions on iodide ions, an intense blue colour appears. This colour disappears as soon as the iodine is consumed by the thiosulphate ions. Thus, the end-point can easily be tracked and the rest is the stoichiometric calculation only.



$$Cu^{2+} + 2I^{\ominus} \longrightarrow \overset{+1}{Cu_2}I_2 + I_2$$



Starch

$$S_2O_3^{2-} + I_2 \longrightarrow S_4O_6^{2-} + I^{\ominus}$$

CuI

CuI

CuI

Cu_2I_2

$MnO_4^{\ominus} + C_2O_4^{2-} \longrightarrow Mn^{2+} + CO_2$

Gram Equivalent amount of $MnO_4^{\ominus}$ = Gram Equivalent amount of Oxalic acid.

Normality = $\dfrac{\text{Gram Equivalent}}{(\text{volume in L})}$

Buffer

Class

Question

Which of the following is not an example of redox reaction ?

(A) $CuO + H_2 \rightarrow Cu + H_2O$

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

(C) $2K + F_2 \rightarrow 2KF$

(D) $BaCl_2 + H_2SO_4 \rightarrow BaSO_4 \downarrow + 2HCl$

→ Not a redox rxn

Question

The more positive the value of E^{Θ}, the greater is the tendency of the species to get reduced. Using the standard electrode potential of redox couples given below find out which of the following is the strongest oxidising agent.

E^{Θ} values: $Fe^{3+}/Fe^{2+} = +0.77$; $I_2(s)/I^- = +0.54$;
$Cu^{2+}/Cu = +0.34$; $Ag^+/Ag = +0.80V$

(A) Fe^{3+}

(B) $I_2(s)$

(C) Cu^{2+}

(D) Ag^+

Question

$E^{\ominus}$ values of some redox couples are given below. On the basis of these values choose the correct option.

$E^{\ominus}$ values: $Br_2/Br^- = +1.90$; $Ag^+/Ag(s) = +0.80$

$Cu^{2+}/Cu(s) = +0.34$; $I_2(s)/I^- = +0.54$

(A) Cu will reduce Br^-

(B) Cu will reduce Ag

(C) Cu will reduce I^-

(D) Cu will reduce Br_2

Question

Using the standard electrode potential, find out the pair between which redox reaction is not feasible.

$E^{\ominus}$ values: $Fe^{3+}/Fe^{2+} = +0.77$; $I_2/I^- = +0.54$;

$Cu^{2+}/Cu = +0.34$; $Ag^+/Ag = +0.80$ V

A. Fe^{3+} and I^-

B. Ag^+ and Cu

C. Fe^{3+} and Cu

D. Ag and Fe^{3+}

Question

Thiosulphate reacts differently with iodine and bromine in the reactions given below:

$$2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$$

$$S_2O_3^{2-} + 2Br_2 + 5H_2O \rightarrow 2SO_4^{2-} + 4Br^- + 10^+$$

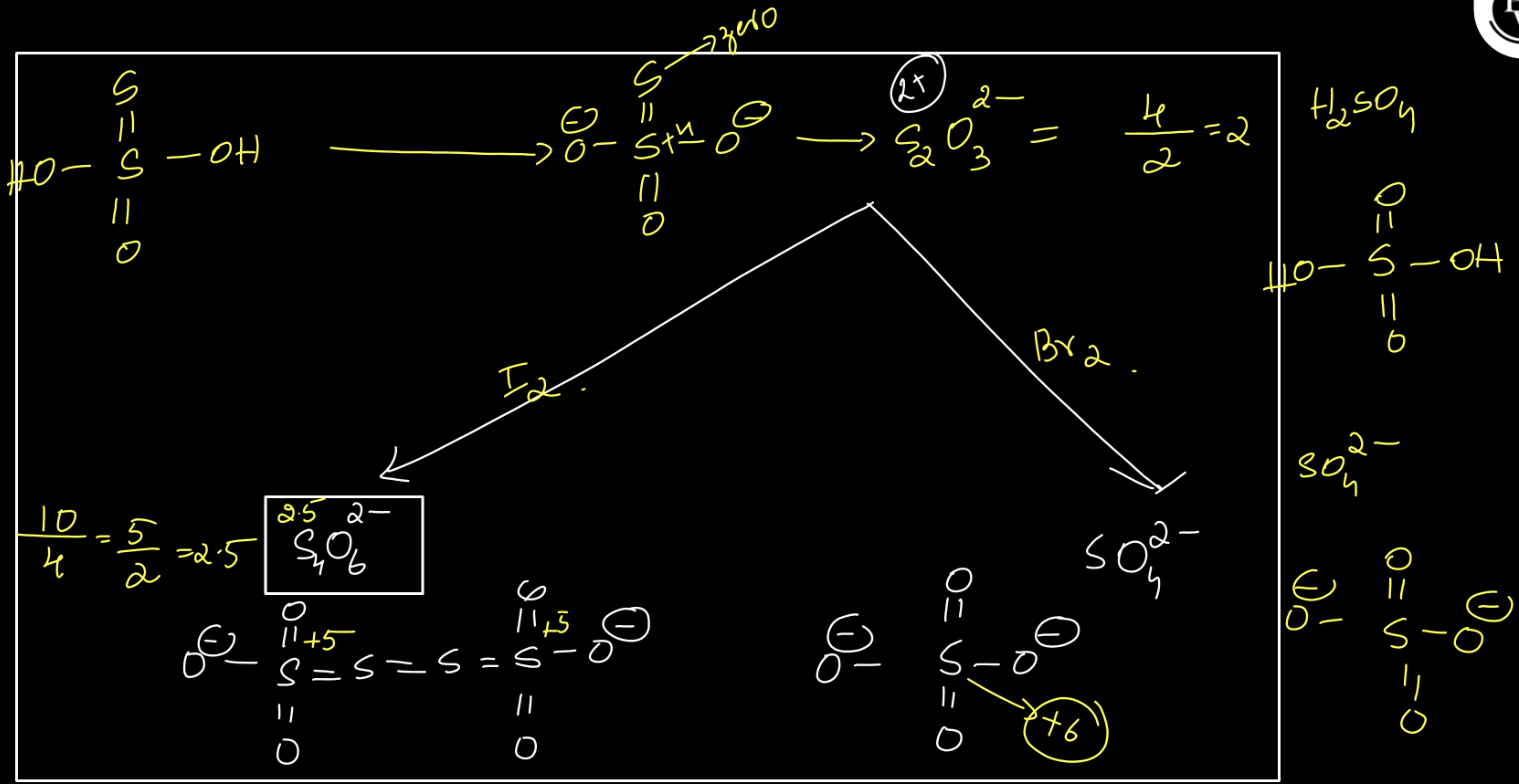
Which of the following statements justifies the above dual behaviour of thiosulphate?

A Bromine is a stronger oxidant than iodine.

B Bromine is a weaker oxidant than iodine.

C Thiosulphate undergoes oxidation by bromine and reduction by iodine in these reactions.

D Bromine undergoes oxidation and iodine undergoes reduction in these reactions.


zero
2+
S2O3 2−
= 4/2 = 2
H2SO4
I2
Br2
10/4 = 5/2 = 2.5
2.5
S4O6 2−
+5
+5
SO4 2−
+6
SO4 2−

Question

Which of the following arrangements represent increasing oxidation number of the central atom?

(A) $CrO_2^-, ClO_3^-, CrO_4^{2-}, MnO_4^-$

(B) $ClO_3^-, CrO_4^{2-}, MnO_4^-, CrO_2^-$

(C) $CrO_2^-, ClO_3^-, MnO_4^-, CrO_4^{2-}$

(D) $CrO_4^{2-}, MnO_4^-, CrO_2^-, ClO_3^-$

$MnO_4^- \rightarrow x + (-2\times4) = -1$

$x = +7$

$CrO_4^{2-} \rightarrow x + (-2\times4) = -2$

$x = +6$

$ClO_3^- \rightarrow x + (-2\times3) = -1$

$x - 6 = -1$

$x = +5$

CrO_2^-

$x - 4 = -1$

$x = +3$

Thank You