FREE NOTES · NCERT + JEE/NEET LEVEL

Redox Reactions

Classical and electronic ideas of oxidation and reduction, oxidation number rules, types of redox reactions, balancing by oxidation number and half-reaction methods, redox titrations and an introduction to electrode processes.

01 Classical definitions
02 Electron transfer view
03 Oxidation number (ON)
04 Types of redox reactions
05 Balancing: oxidation number method
06 Balancing
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01 Classical definitions

Oxidation	Reduction
Addition of oxygen / electronegative element	Removal of oxygen / electronegative element
Removal of hydrogen / electropositive element	Addition of hydrogen / electropositive element

$2Mg + O_2 \rightarrow 2MgO$ (Mg oxidised). $CuO + H_2 \rightarrow Cu + H_2O$ (CuO reduced, H_2 oxidised). Both happen together ⇒ **redox**.

02 Electron transfer view

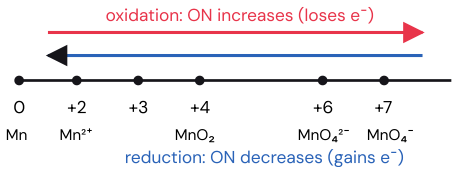
OIL RIG

Oxidation **I**s **L**oss of electrons; **R**eduction **I**s **G**ain. The species that is oxidised is the **reducing agent**; the species reduced is the **oxidising agent**.

$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$: Zn loses $2e^-$ (oxidised, reducing agent); Cu^{2+} gains $2e^-$ (reduced, oxidising agent). Zn displaces Cu because Zn gives up electrons more easily (electrochemical series).

03 Oxidation number (ON)

1. Free element: 0 (O_2, Na, P_4, S_8).
2. Monoatomic ion: its charge (Na^+ +1, Cl^- −1).
3. O is −2, except peroxides (−1: H_2O_2, Na_2O_2), superoxides (−½: KO_2), OF_2 (+2), O_2F_2 (+1).
4. H is +1, except metal hydrides (−1: NaH, CaH_2).
5. F always −1. Other halogens −1 except with O or a more electronegative halogen.
6. Alkali metals +1, alkaline earth +2.
7. Sum of ONs = 0 in a molecule, = charge in an ion.



Oxidation states of Mn: up = oxidation, down = reduction

Watch out: in CrO_5 (butterfly), Cr is +6 (four O are peroxide, −1). In Fe_3O_4, average ON of Fe = +8/3. In $S_4O_6^{2-}$ average +2.5 (two S are +5, two are 0).

04 Types of redox reactions

Type	Example
Combination	$C + O_2 \rightarrow CO_2$
Decomposition	$2KClO_3 \rightarrow 2KCl + 3O_2$
Metal displacement	$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$
Non-metal displacement	$2Na + 2H_2O \rightarrow 2NaOH + H_2$; $Cl_2 + 2KBr \rightarrow 2KCl + Br_2$
Disproportionation (same element oxidised and reduced)	$2H_2O_2 \rightarrow 2H_2O + O_2$; $Cl_2 + 2OH^- \rightarrow ClO^- + Cl^- + H_2O$; $P_4 + NaOH \rightarrow PH_3 + NaH_2PO_2$
Comproportionation (reverse)	$Ag^{2+} + Ag \rightarrow 2Ag^+$

F_2 cannot disproportionate (it can only be reduced: no positive ON). ClO_4^- cannot disproportionate (Cl already at max +7).

05 Balancing: oxidation number method

1. Write skeleton equation; find atoms whose ON changes.
2. Find increase and decrease of ON per formula unit.
3. Multiply to make total increase = total decrease.
4. Balance other atoms except O and H.
5. Balance O with H_2O, then H with H^+ (acid) or with H_2O and OH^- (base). Check charges.

Example: MnO_4^- + Fe^{2+} → Mn^{2+} + Fe^{3+} (acid). Mn: +7 → +2 (−5); Fe: +2 → +3 (+1). Multiply Fe by 5 ⇒ $MnO_4^- + 5Fe^{2+} + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$.

06 Balancing: half-reaction (ion-electron) method

1. Split into oxidation and reduction half-reactions.
2. Balance atoms other than O, H; then O with H_2O, H with H^+.
3. Balance charge with electrons.
4. In basic medium, add OH^- to both sides to cancel H^+.
5. Multiply halves so electrons cancel; add.

Example: $Cr_2O_7^{2-}$ + Fe^{2+} (acid). Reduction: $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$. Oxidation: $Fe^{2+} \rightarrow Fe^{3+} + e^-$ (× 6). Sum: $Cr_2O_7^{2-} + 6Fe^{2+} + 14H^+ \rightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$.

07 Redox titrations

Titrant	Indicator
$KMnO_4$ (acid)	Self indicator: first permanent pink colour
$K_2Cr_2O_7$	Diphenylamine (blue end point)
I_2 / $Na_2S_2O_3$ (iodometry)	Starch (blue colour disappears)

Iodometry: $2Cu^{2+} + 4I^- \rightarrow Cu_2I_2 + I_2$, then $I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$.

08 Redox and electrode processes

If Zn and $CuSO_4$ are kept in separate beakers joined by a wire and salt bridge, the same redox reaction drives an electric current: this is a **galvanic cell** (Daniell cell). Each half is a **redox couple** (Zn^{2+}/Zn, Cu^{2+}/Cu). Standard electrode potential E° measures the tendency to be reduced; more positive E° ⇒ stronger oxidising agent. (Full detail in Class 12 Electrochemistry.)

09 JEE Extra + Tricks

JEE LEVEL

- Oxidation number tricks: in peroxides O = −1 (H_2O_2, Na_2O_2), superoxides −½ (KO_2), OF_2 O = +2; in CrO_5 and H_2SO_5 use structure (peroxide bonds).
- Average O.N. of S in $Na_2S_4O_6$ (tetrathionate) = +2.5 (actual 0 and +5); Fe_3O_4 Fe = +8/3.
- n-factor in redox = change in O.N. × number of atoms per molecule; equivalents of oxidant = equivalents of reductant.
- $KMnO_4$ n-factor: acidic 5 (→ Mn^{2+}), neutral 3 (→ MnO_2), strongly basic 1 (→ MnO_4^{2-}). $K_2Cr_2O_7$ (acidic): 6.
- Disproportionation: same element oxidised and reduced (Cl_2 in NaOH, $H_2O_2 \rightarrow H_2O + O_2$).

10 Solved questions

Q1. ON of S in H_2SO_4, of Cr in $K_2Cr_2O_7$, of P in H_3PO_2.

S: 2 + x − 8 = 0 ⇒ **+6**. Cr: 2 + 2x − 14 = 0 ⇒ **+6**. P: 3 + x − 4 = 0 ⇒ **+1**.

Q2. Identify oxidising and reducing agents: $H_2S + Cl_2 \rightarrow 2HCl + S$.

S: −2 → 0 (oxidised) ⇒ H_2S is the reducing agent. Cl: 0 → −1 ⇒ Cl_2 is the oxidising agent.

Q3. Balance in basic medium: $MnO_4^- + I^- \rightarrow MnO_2 + I_2$.

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

11 Mistakes & practice

- Calling the oxidised substance the oxidising agent.
- Giving O as −2 in H_2O_2 or KO_2.
- Forgetting to convert H^+ to OH^- in basic medium.
- Thinking ON must be a whole number (averages can be fractions).

1. ON of O in OF_2?
2. Is $2NO_2 \rightarrow N_2O_4$ a redox reaction?
3. Name the disproportionation in $3HNO_2 \rightarrow HNO_3 + 2NO + H_2O$.
4. Indicator in $KMnO_4$ titration?

ANSWERS

1) +2 · 2) No, N stays +4 · 3) N: +3 → +5 and +2 · 4) $KMnO_4$ itself

12 Quick revision: 5 minutes before the exam

BASICS

- OIL RIG.
- Oxidised = reducing agent.
- ON ↑ oxidation, ↓ reduction.

ON RULES

- O −2 (peroxide −1, OF_2 +2).
- H +1 (hydride −1).
- F always −1.

TYPES

- Combination, decomposition, displacement.
- Disproportionation: same element up and down.

BALANCING

- Equalise ON change.
- O with H_2O, H with H^+.
- Basic: add OH^-.

Have a doubt? Ask the AI Doubt Solver at padhaiadda.netlify.app, or take the mock test on this chapter.