FREE EDITION · NOTES + 3 SAMPLE MCQS

CUET · CHEMISTRY · CLASS XI · CODE 306

Redox Reactions

CUET unit: Redox Reactions

By **UniDrill** · NCERT-grounded study material

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📌 Snapshot

- Builds the three-tier conceptualisation of redox — classical (oxygen/hydrogen transfer), electronic (electron transfer) and oxidation-number based — and shows that all three converge on the same set of reactions.
- Trains the student in assigning oxidation numbers using a set of six rules, including the awkward cases of peroxides, superoxides, oxygen-fluorine compounds and mixed oxides (Fe_3O_4, Mn_3O_4, Pb_3O_4) where fractional ON appears.
- Classifies redox reactions into four operational types — combination, decomposition, displacement (metal/non-metal) and disproportionation — with comproportionation appearing implicitly through reverse-disproportionation logic.
- Develops both balancing techniques NTA loves to test: oxidation-number method and half-reaction (ion-electron) method, in acidic and basic media.
- Closes by linking redox to electrode processes — Daniell cell, salt bridge, standard electrode potential, the electrochemical series in Table 7.1 (Li at top of reducing power, F_2 at top of oxidising power) — and to redox titrations ($KMnO_4$ self-indicator, $K_2Cr_2O_7$ with diphenylamine, iodine-thiosulphate with starch).

📖 Detailed Notes

2.1 Core concepts

- **Classical idea of oxidation** was originally addition of oxygen ($2Mg + O_2 \rightarrow 2MgO$; $S + O_2 \rightarrow SO_2$; $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$), later widened to addition of any electronegative element OR removal of hydrogen/electropositive element from a substance (NCERT §7.1, p. 235–236).
- **Classical idea of reduction** is the mirror image — removal of oxygen/electronegative element ($2HgO \rightarrow 2Hg + O_2$; $2FeCl_3 + H_2 \rightarrow 2FeCl_2 + 2HCl$) or addition of hydrogen/electropositive element ($CH_2{=}CH_2 + H_2 \rightarrow H_3C\text{-}CH_3$; $2HgCl_2 + SnCl_2 \rightarrow Hg_2Cl_2 + SnCl_4$). Oxidation and reduction always occur simultaneously — hence the word "redox" (NCERT §7.1, p. 236).
- **Electron-transfer concept:** In $2Na + Cl_2 \rightarrow 2NaCl$, the half reactions $2Na \rightarrow 2Na^+ + 2e^–$ (oxidation = loss of electrons) and $Cl_2 + 2e^– \rightarrow 2Cl^–$ (reduction = gain of

electrons) define oxidation as electron loss, reduction as electron gain. Oxidising agent = electron acceptor; reducing agent = electron donor (NCERT §7.2, p. 237).

- **Competitive electron transfer** is illustrated by $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$ (equilibrium lies far to the right) and $Cu + 2Ag^{+} \rightarrow Cu^{2+} + 2Ag$, while $Co + Ni^{2+} \rightleftharpoons Co^{2+} + Ni$ reaches an intermediate equilibrium — giving the order of reducing tendency Zn > Cu > Ag, the embryo of the electrochemical/activity series (NCERT §7.2.1, p. 238–239).
- **Oxidation number** assumes complete transfer of the bonding pair to the more electronegative atom — a book-keeping device. Six rules govern its assignment: (1) ON of an element in its free state is zero (H_2, O_2, Cl_2, P_4, S_8, Na, Mg, Al); (2) for monoatomic ions ON equals the charge ($Na^{+} = +1$, $Mg^{2+} = +2$, $Cl^{-} = -1$); (3) ON of O is normally –2, but –1 in peroxides (H_2O_2, Na_2O_2), –½ in superoxides (KO_2, RbO_2), and +2 in OF_2, +1 in O_2F_2; (4) ON of H is +1, but –1 in metal hydrides like LiH, NaH, CaH_2; (5) F is always –1; other halogens are –1 with metals but positive when bonded to O; (6) algebraic sum of ON = 0 in a neutral compound and equals the charge in a polyatomic ion (NCERT §7.3, p. 239–240).
- **Stock notation** writes the oxidation state of a metal as a Roman numeral after its symbol — Au(I)Cl, $Au(III)Cl_3$, $Sn(II)Cl_2$, $Sn(IV)Cl_4$, Mn(II)O, $Mn(IV)O_2$, Fe(II)O, $Fe_2(III)O_3$ — useful for distinguishing reduced and oxidised forms of the same metal (NCERT §7.3, p. 241).
- **Oxidation number based definitions:** oxidation = increase in ON; reduction = decrease in ON; oxidant = species that increases ON of another; reductant = species that decreases ON of another; redox reaction = reaction with ON change of interacting species (NCERT §7.3, p. 241).
- **Fractional oxidation numbers** in C_3O_2 (C = +4/3), Br_3O_8 (Br = +16/3) and $Na_2S_4O_6$ (S = +5/2) are averages — structurally the atoms exist in different whole-number states (in C_3O_2 two terminal C atoms are +2 and the middle C is 0; in $S_4O_6^{2-}$ two terminal S are +5 and two middle S are 0). Fe_3O_4, Mn_3O_4 and Pb_3O_4 are mixed oxides where the same fractional-ON paradox appears (Pb_3O_4 is $2PbO \cdot PbO_2$; Fe_3O_4 is $FeO \cdot Fe_2O_3$) (NCERT §7.3 box, p. 244–245).
- **Combination redox reactions** A + B → C require at least one of A, B to be elemental — e.g. $C + O_2 \rightarrow CO_2$; $3Mg + N_2 \rightarrow Mg_3N_2$; $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$. All combustion of dioxygen is in this class (NCERT §7.3.1, p. 242).
- **Decomposition redox reactions** are the reverse — a compound breaks down with at least one product in elemental form: $2H_2O \rightarrow 2H_2 + O_2$; $2NaH \rightarrow 2Na + H_2$; $2KClO_3 \rightarrow 2KCl + 3O_2$. Not every decomposition is redox — $CaCO_3 \rightarrow CaO + CO_2$ involves no ON change (NCERT §7.3.1, p. 242).
- **Displacement reactions** of form X + YZ → XZ + Y come in two flavours. **Metal displacement** — $CuSO_4 + Zn \rightarrow ZnSO_4 + Cu$; $V_2O_5 + 5Ca \rightarrow 2V + 5CaO$; $TiCl_4 + 2Mg \rightarrow Ti + 2MgCl_2$; $Cr_2O_3 + 2Al \rightarrow Al_2O_3 + 2Cr$ — drives industrial metallurgy (thermite process etc.) (NCERT §7.3.1, p. 242–243).

- **Non-metal displacement** is dominated by hydrogen displacement: alkali metals and Ca/Sr/Ba displace H_2 from cold water ($2Na + 2H_2O \rightarrow 2NaOH + H_2$); Mg and Fe need steam ($Mg + 2H_2O \rightarrow Mg(OH)_2 + H_2$; $2Fe + 3H_2O \rightarrow Fe_2O_3 + 3H_2$); less active metals (Zn, Mg, Fe, Cd, Sn) liberate H_2 from acids; Ag and Au don't react with HCl. Halogen displacement runs F > Cl > Br > I ($Cl_2 + 2KBr \rightarrow 2KCl + Br_2$; $Cl_2 + 2KI \rightarrow 2KCl + I_2$; $Br_2 + 2I^- \rightarrow 2Br^- + I_2$) — basis of the "layer test" (NCERT §7.3.1, p. 243).
- **Disproportionation** — one species in an intermediate oxidation state is simultaneously oxidised and reduced. The element must have at least three accessible oxidation states. Examples: $2H_2O_2 \rightarrow 2H_2O + O_2$ (O at $-1 \rightarrow -2$ and 0); $P_4 + 3OH^- + 3H_2O \rightarrow PH_3 + 3H_2PO_2^-$ (P: $0 \rightarrow -3$ and $+1$); $S_8 + 12OH^- \rightarrow 4S^{2-} + 2S_2O_3^{2-} + 6H_2O$ (S: $0 \rightarrow -2$ and $+2$); $Cl_2 + 2OH^- \rightarrow ClO^- + Cl^- + H_2O$ — basis of household bleach (Cl: $0 \rightarrow +1$ and -1) (NCERT §7.3.1, p. 244).
- **Fluorine deviation:** $F_2 + 2OH^- \rightarrow 2F^- + OF_2 + H_2O$ — F has no positive oxidation state available, so it does **not** disproportionate. Among ClO^-, ClO_2^-, ClO_3^-, ClO_4^-, only ClO_4^- does not disproportionate because Cl is already in its highest state +7 (NCERT §7.3.1, p. 244).
- **Balancing — oxidation number method:** (1) write skeletal equation; (2) assign ON to identify oxidant and reductant; (3) make ON-increase equal to ON-decrease by suitable multiplication; (4) balance ionic charges with H^+ (acidic medium) or OH^- (basic medium); (5) balance H atoms with H_2O — final cross-check that O atoms balance (NCERT §7.3.2, p. 246).
- **Balancing — half-reaction (ion-electron) method:** (1) write unbalanced ionic equation; (2) split into oxidation and reduction halves; (3) balance atoms other than O and H first; (4) in acidic medium balance O with H_2O and H with H^+; (5) balance charge with electrons; (6) equalise electrons across the two halves and add; (7) verify. **For basic medium**, balance as if acidic, then add equal OH^- to both sides to neutralise H^+; combine $H^+ + OH^- \rightarrow H_2O$ (NCERT §7.3.2, p. 246–249). Worked examples in chapter: $Cr_2O_7^{2-} + 3SO_3^{2-} + 8H^+ \rightarrow 2Cr^{3+} + 3SO_4^{2-} + 4H_2O$; $2MnO_4^- + Br^- + H_2O \rightarrow 2MnO_2 + BrO_3^- + 2OH^-$; $6Fe^{2+} + Cr_2O_7^{2-} + 14H^+ \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O$; $6I^- + 2MnO_4^- + 4H_2O \rightarrow 3I_2 + 2MnO_2 + 8OH^-$.
- **Redox titrations** rely on indicator-controlled colour change. (i) **$KMnO_4$** acts as its own indicator — pink persists at MnO_4^- ~10^{-6} mol L^{-1} just past equivalence. (ii) **$K_2Cr_2O_7$** is not a self-indicator; **diphenylamine** is oxidised immediately past the equivalence point giving intense blue. (iii) **Iodometric** titrations use the starch–iodine deep-blue with Cu^{2+} liberating I_2 from KI, the I_2 then consumed by thiosulphate via $I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$ (NCERT §7.3.3, p. 249).
- **Redox in electrochemical cells:** the same $Zn + CuSO_4$ reaction, when zinc and copper rods are separated into two beakers connected by a salt bridge (U-tube of KCl/NH_4NO_3 in agar) and joined by an external wire, becomes the **Daniell cell**. Electrons travel externally from Zn anode (oxidation) to Cu cathode (reduction); ions migrate via the salt bridge (NCERT §7.4, p. 250).

- **Redox couple** = oxidised/reduced form of the same species (Zn2+/Zn, Cu2+/Cu) — oxidised form is written first. The **electrode potential** measures the tendency of the couple to remain in oxidised or reduced form. At unit activity and 298 K it is the **standard electrode potential E°** , with E°(H+/H2) = 0.00 V by convention. Negative E° means stronger reducing agent than H+/H2; positive E° means weaker reducing agent than H+/H2 (NCERT §7.4, p. 250–251).
- **Electrochemical series (Table 7.1):** in descending oxidising strength — F2 (+2.87), Co3+ (+1.81), H2O2 (+1.78), MnO4– (+1.51 acidic), Au3+ (+1.40), Cl2 (+1.36), Cr2O72– (+1.33), O2 (+1.23), MnO2 (+1.23), Br2 (+1.09), Ag+ (+0.80), Fe3+/Fe2+ (+0.77), I2 (+0.54), Cu2+ (+0.34), H+/H2 (0.00), Pb2+ (–0.13), Sn2+ (–0.14), Ni2+ (–0.25), Fe2+ (–0.44), Cr3+ (–0.74), Zn2+ (–0.76), Al3+ (–1.66), Mg2+ (–2.36), Na+ (–2.71), Ca2+ (–2.87), K+ (–2.93), Li+ (–3.05). Reducing strength increases down the table; oxidising strength increases up (NCERT Table 7.1, p. 251).

2.2 Definitions to memorise

Term	Definition	Page
Oxidation (classical)	Addition of oxygen/electronegative element OR removal of hydrogen/electropositive element from a substance	236
Reduction (classical)	Removal of oxygen/electronegative element OR addition of hydrogen/electropositive element to a substance	236
Oxidation (electronic)	Loss of electron(s) by any species	237
Reduction (electronic)	Gain of electron(s) by any species	237
Oxidising agent	Acceptor of electron(s); increases the oxidation number of the other species	237, 241
Reducing agent	Donor of electron(s); lowers the oxidation number of the other species	237, 241
Oxidation number	Charge an atom would carry if all bond pairs in covalent bonds were assigned to the more electronegative atom	239
Oxidation (by ON)	An increase in the oxidation number of the element in the given substance	241
Reduction (by ON)	A decrease in the oxidation number of the element in the given substance	241
Stock notation	Oxidation state shown as a Roman numeral in parentheses after the metal symbol, e.g. Au(III)Cl3	241
Combination reaction	A + B → C where at least one reactant is in elemental form (e.g. C + O2 → CO2)	242
Decomposition reaction	Compound breaks down such that at least one product is in elemental form (e.g. 2H2O → 2H2 + O2)	242
Displacement reaction		242

Term	Definition	Page
	X + YZ → XZ + Y — an atom/ion in a compound is replaced by an atom/ion of another element	
Disproportionation reaction	A redox reaction in which an element in an intermediate oxidation state is simultaneously oxidised and reduced	244
Redox couple	The oxidised and reduced forms of the same substance taken together (e.g. Zn2+/Zn, Cu2+/Cu)	250
Electrode potential	Potential associated with each electrode; a measure of the tendency of the active species to stay in oxidised/reduced form	250
Standard electrode potential (E°)	Electrode potential when all species have unit activity, gases at 1 atm, at 298 K; E°(H+/H2) = 0.00 V by convention	250
Salt bridge	U-tube of KCl or NH4NO3 in agar that completes the circuit between the two half-cells without letting solutions mix	250

2.3 Diagrams / processes to remember

- **Fig. 7.1 — Zn strip in aqueous Cu(NO3)2:** zinc strip becomes coated with reddish copper; blue colour fades as Cu2+ is reduced (NCERT §7.2.1, p. 238).
- **Fig. 7.2 — Cu strip in aqueous AgNO3:** solution turns blue (Cu → Cu2+) and silver deposits on the strip (NCERT §7.2.1, p. 239).
- **Fig. 7.3 — Daniell cell:** Zn rod in ZnSO4 + Cu rod in CuSO4 connected by a salt bridge and external wire with ammeter. Electrons flow Zn → Cu through wire; conventional current is opposite (NCERT §7.4, p. 250).
- **Highest-oxidation-number table (p. 241):** Na(+1), Mg(+2), Al(+3), Si(+4), P(+5), S(+6), Cl(+7) — highest ON of a representative element generally equals its group number (or group number – 10 for groups 13–17).
- **Table 7.1 — Standard electrode potentials at 298 K:** memorise that F2 sits at the top (+2.87 V, strongest oxidant) and Li at the bottom (–3.05 V, strongest reductant); H+/H2 is the zero (NCERT Table 7.1, p. 251).
- **Structural pictures of "fractional ON" species** (p. 245): C3O2 = O=C=C*=C=O (terminal C +2, middle C 0); S4O62– (terminal S +5, middle S 0); Br3O8 (terminal Br +6, middle Br +4). These show fractional ON is an average artefact.

2.4 Common confusions / NTA trap points

- "Every decomposition reaction is a redox reaction" — false. CaCO3 → CaO + CO2 has no change in oxidation number; only decompositions where at least one product is elemental are redox (e.g. 2KClO3 → 2KCl + 3O2 is redox, 2NaH → 2Na + H2 is redox, but CaCO3 decomposition is not) (NCERT §7.3.1, p. 242).

- ON of oxygen is **not always –2**: in peroxides it is –1 (H2O2, Na2O2), in superoxides it is –½ (KO2, RbO2), in OF2 it is +2 and in O2F2 it is +1 — NTA loves to put H2O2 or OF2 in the stem (NCERT §7.3 rule 3, p. 240).
- ON of hydrogen is **+1 except in metal hydrides** (LiH, NaH, CaH2, NaBH4 — where it's –1). Students who blindly write +1 for H in NaH will miss the reduction step in 2NaH → 2Na + H2 (NCERT §7.3 rule 4, p. 240).
- **Fluorine never shows a positive oxidation state and therefore never disproportionates.** Among ClO–, ClO2–, ClO3–, ClO4–, the one that does **not** disproportionate is ClO4– (Cl already in +7, no higher state available) — NTA frequently sets this as the trap (NCERT §7.3.1 + Problem 7.5, p. 244).
- KMnO4 acts as its **own indicator** because of its intense purple colour; K2Cr2O7 does **not** — it needs diphenylamine. Mixing these up is a classic trap (NCERT §7.3.3, p. 249).
- The half-reaction method in basic medium does **not** add OH– directly to balance O — you balance as in acid first with H+/H2O, then neutralise each H+ with an OH– on both sides, combining the new H+ + OH– into H2O. Skipping this routine gives the wrong stoichiometry.

🎯 Practice MCQs

Q1. The oxidation number of sulphur in the tetrathionate ion S4O62– is, on average:

A. +2

B. +2.5

C. +3

D. +5

Q2. Which of the following decomposition reactions is **not** a redox reaction?

A. 2H2O(l) → 2H2(g) + O2(g)

B. 2NaH(s) → 2Na(s) + H2(g)

C. 2KClO3(s) → 2KCl(s) + 3O2(g)

D. CaCO3(s) → CaO(s) + CO2(g)

Q3. The oxidation number of iron in the mixed oxide Fe3O4 is:

- **A.** +2
- **B.** +3
- **C.** +8/3
- **D.** +4

📊 PYQ Alignment

Redox Reactions is one of the most frequently tested Class XI Chemistry chapters in CUET (UG), typically yielding 5–8 MCQs per paper across the 2023–2025 cycle. The recurring question patterns are: assigning oxidation number to underlined atoms (especially in H2SO5, Cr2O72–, NO3–, Fe3O4, NaH, KO2, OF2 and S4O62–), identifying the type of redox reaction from a given equation (combination/decomposition/displacement/disproportionation), spotting the oxidising and reducing agents in a balanced equation, predicting feasibility from standard electrode potentials in Table 7.1, and short stoichiometric balancing problems in acidic or basic medium using the ion-electron method.

CUET 2025 — Actual PYQs from this chapter

Q.16 (CUET 2025) Why is HCl not used to make medium acidic in oxidation reactions of $KMnO_4$?

- A) $KMnO_4$ oxidises HCl B) $KMnO_4$ reduces HCl C) HCl acts as oxidising agent D) $KMnO_4$ acts as reducing agent

Tests: Redox Reactions Answer: Not in extracted key

CUET 2024 — Actual PYQs from this chapter

Q.4 (CUET 2024) The oxidation number of Co in the complex [Co(H2NCH2CH2NH2)3]2(SO4)3[Co(H_2NCH_2CH_2N H_2)_3]_2(SO_4)_3[Co(H2 NCH2 CH2 NH2)3]2 (SO4)3 is:

- A) 3 B) 4 C) 2 D) 5

Tests: Redox Reactions Answer: Not in extracted key

CUET 2023 — Actual PYQs from this chapter

Q.17 (CUET 2023) Identify X with oxidation state of halogen: $3Cl_2 + 6NaOH \rightarrow 5NaCl + X + 3H_2O$ (hot & conc.)

- A) $NaClO_2$, +3 B) $NaClO_3$, +5 C) $NaClO_2$, +7 D) $NaClO_3$, +1

Tests: Disproportionation of Cl2/NaOH **Answer:** Not in extracted key

Q.15 (CUET 2023) In cyanide extraction of silver the oxidising and reducing agents are respectively:

- A) O_2 and CO B) O_2 and Zn dust C) H_2O and Zn dust D) H_2O and NaCN

Tests: Redox in metallurgical extraction (cyanide process) **Answer:** Not in extracted key

Q.18 (CUET 2023) Match List-I with List-II. List-I List-II (A) Electrolytic reduction (I) Iron (B) Bessemerisation (II) Zinc (C) Smelting (III) Aluminium (D) Reduction from oxide (IV) Copper Options:

- A) A-III, B-I, C-II, D-IV B) A-II, B-III, C-IV, D-I C) A-III, B-IV, C-I, D-II D) A-IV, B-II, C-III, D-I

Tests: Redox methods in metallurgy (electrolytic reduction/Bessemerisation) **Answer:** Not in extracted key