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REDOX REACTIONS

Class 11 Chemistry • Oxidation Number & n-Factor • NEET / JEE Notes

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Topics Covered

1. Redox reactions — oxidation & reduction (classical and electronic concepts)
2. Oxidation number (ON) — rules, worked examples, coordinate bond cases
3. Minimum & maximum oxidation states; peroxide-linkage traps (H_2SO_5, $H_2S_2O_8$, CrO_5, K_3CrO_8)
4. Oxidising / reducing behaviour from ON; average ON (C_3O_2, Br_3O_8 — PYQ)
5. n-factor — acids, bases, salts, ions, elements, reactions
6. Added: equivalent weight, normality, law of equivalence & titrations, balancing redox equations, practice MCQs

REDOX REACTIONS

Reactions in which **oxidation and reduction occur simultaneously**.

Oxidation	Reduction
Loss of e^-	Gain of e^-
Removal of H_2	Addition of H_2
Addition of O_2	Removal of O_2
Oxidation number ↑	Oxidation number ↓

Oxidation number depends on electronegativity (EN)

In a bond A–B, the **more EN** atom takes the shared electrons ⇒ **–ve** ON; the **less EN** atom gets **+ve** ON (e.g. Na^+Cl^-).

Majority cases: **O = –2**, **H = +1**, **F = –1** (always), alkali metals +1, alkaline-earth metals +2; sum of ON = charge on species.

OXIDATION NUMBER — WORKED EXAMPLES

#	Compound	Working	Result
1	HF	H less EN than F	H = +1, F = –1
2	H_2	ΔEN = 0 (same element)	H = 0
3	H_2O_2	Peroxide linkage (O–O)	H = +1, O = –1 (each)
4	O_2F_2	F more EN than O	F = –1, O = +1 (each)

#	Compound	Working	Result
5	KO_2	Superoxide: 2[O] = –1	K = +1, O = –½ (each)
6	H_2SO_4	2(+1) + x + 4(–2) = 0	S = +6
7	$HClO_4$	1 + x + 4(–2) = 0	Cl = +7
8	NO_3^-	x + 3(–2) = –1	N = +5
9	$K_4[Fe(CN)_6]$	4(+1) + x + 6(–1) = 0	Fe = +2
10	$[Fe(H_2O)_5NO]SO_4$ (brown ring, test for NO_3^-)	x + 5(0) + (+1) + (–2) = 0 (NO^+ = nitrosonium)	Fe = +1

▸ Coordinate Bond — VVV IMP

Case	Rule	Example
Case 1: EN(donor) > EN(acceptor)	**Ignore** the coordinate bond (donor → 0, acceptor → 0)	$H_3N \rightarrow BF_3$ type dative bonds
Case 2: EN(donor) < EN(acceptor)	Donor → **+2**, acceptor → **–2**	N → O coordinate bond (as in NO_3^-, NO_2 structures)

MINIMUM & MAXIMUM OXIDATION STATES

Group	Valence e^-	Min OS	Max OS
Carbon family (14)	4	–4	+4
Nitrogen family (15)	5	–3	+5
Oxygen family (16)	6	–2	+6
Halogens (17)	7	–1	+7

- Max OS: **Mn = +7**, **Cr = +6**; in the whole periodic table the maximum is **+8 → Os, Ru**.
- **C^{+5} and O^{+6} are not possible** — valence electrons are the hard ceiling; **S^{+6} is possible** (expanded octet).
- **F is always –1** (never positive): F^{+1}, F^{+2}, F^{+3} are not possible.

PEROXIDE LINKAGE TRAP — FORMULA METHOD FAILS

Compound	Naive formula result	Structure fact	Correct ON
Caro's acid H_2SO_5	x = +8 (wrong)	One O–O peroxide linkage (2 O at –1)	**S = +6**
Marshall's acid $H_2S_2O_8$	x = +7 (wrong)	One O–O linkage joins the two –SO_3H halves	**S = +6** (each)
CrO_5 (butterfly)	x = +10 (wrong)	2 peroxide linkages (4 O at –1) + 1 O (=O) at –2	**Cr = +6**
K_3CrO_8 (crab)	x = +13 (wrong)	4 peroxide linkages — all 8 O at –1	**Cr = +5**

NEET Point / PYQ Alert

Whenever a structure has an **O–O bond**, treat those oxygens as **–1 (peroxide)**, not –2, then solve for the central atom.

Always check the structure before trusting the plug-and-chug formula.

OXIDISING & REDUCING BEHAVIOUR

Oxidation state of element	Behaves as	Examples
Maximum OS	**Only oxidising agent (OA)**	$KMnO_4$ (Mn^{+7}), $K_2Cr_2O_7$ (Cr^{+6}), HNO_3 (N^{+5})
Minimum OS	**Only reducing agent (RA)**	CH_4 (C^{-4}), NH_3 (N^{-3}), H_2S (S^{-2})
Intermediate OS	**Both OA and RA**	HNO_2 (N^{+3}), H_2O_2 (O^{-1}), SO_2 (S^{+4})

AVERAGE OXIDATION STATE — NCERT NAMED COMPOUNDS (2016 & 2018 PYQ)

Compound	Formula method (average)	Structure	Atom-wise ON
C_3O_2 (carbon suboxide)	$3x + 2(-2) = 0 \Rightarrow$ **+4/3**	O=C=C=C=O	+2, 0, +2
Br_3O_8 (tribromo octaoxide)	$3x + 8(-2) = 0 \Rightarrow$ **+16/3**	$O_3Br–BrO_2–BrO_3$	+6, +4, +6

NEET Point / PYQ Alert

When atoms of the same element are in **different environments**, the formula method only gives an **average** ON — draw the structure to get the real value (also: Fe_3O_4 = 8/3; $S_4O_6^{2-}$ = 2.5 average, actual 0, +5).

N-FACTOR (VVV IMP)

Just a number — it can be calculated for **acids, bases, salts, ions (charge), elements and reactions**.

Category	n-factor =
Acids	Number of **H^+** ions given (ionisable H)
Bases	Number of **OH^-** ions given
Salts	Total **+ve or –ve charge**
Ions (charge)	**Magnitude of charge**
Elements	**Valency**
Redox reaction	Σ [change in ON × no. of atoms] (per formula unit of reactant)

Species	n-factor	Species	n-factor
HCl	1	NaOH	1
NaCl	1	H_2SO_4	2
$Ca(OH)_2$	2	$Al_2(SO_4)_3$	6
O^{2-}	2	N (element)	3
Mg	2	Au^{3+}	3
Na	1	Alum $K_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O$	8 (water of crystallisation ignored)

n-Factor of Acids — Only H Attached to O (Ionisable) Counts

Acid	Ionisable H	n-factor (basicity)
H_3PO_4 (phosphoric)	All 3 H on O	3
H_3PO_3 (phosphorous)	1 H bonded directly to P (non-ionisable)	2
H_3PO_2 (hypophosphorous)	2 H bonded directly to P	1
$H_3BO_3 = B(OH)_3$ (orthoboric)	Lewis acid — accepts OH^-: $B(OH)_3 + H_2O \rightarrow [B(OH)_4]^- + H^+$	1

n-Factor via the Balanced Reaction

- $H_3PO_4 + NaOH \rightarrow NaH_2PO_4 + H_2O \Rightarrow$ **n = 1**; $H_3PO_4 + 2NaOH \rightarrow Na_2HPO_4 + 2H_2O \Rightarrow$ **n = 2**.
- $Al(OH)_3 + HCl \rightarrow Al(OH)_2Cl + H_2O \Rightarrow$ **n = 1**.
- The n-factor is the number of H^+/OH^- **actually exchanged in that reaction** — not necessarily the full acidity/basicity.

n-Factor for Redox Reactions

Rule

$$\text{n-factor} = \Sigma\,[\text{change in ON} \times \text{number of atoms}] \quad \text{(of the reactant)}$$

$KMnO_4$ (Mn^{+7}) in	Product	n-factor
Acidic medium	Mn^{2+}	5 × 1 = **5**
Neutral / weakly basic	MnO_2 (Mn^{4+})	3 × 1 = **3**
Strongly basic	MnO_4^{2-} (Mn^{6+})	1 × 1 = **1**

Reaction	n-factor
$K_2Cr_2O_7$ (Cr^{+6}) —acidic→ Cr^{3+}	3 × 2 = **6**
FeC_2O_4 (ferrous oxalate: Fe^{2+}, C^{+3}) → $Fe^{3+} + CO_2$ (C^{+4})	(1×1) + (1×2) = **3**
$Fe_2(C_2O_4)_3$ (Fe^{3+}, C^{+3}) → CO_2 (C^{+4}); Fe unchanged	1 × 6 C atoms = **6**

Reaction	n-factor
$FeSO_3$ (Fe^{2+}, S^{+4}) → Fe^{3+} + SO_4^{2-} (S^{+6})	(1×1) + (2×1) = **3**
$FeSO_4$ (Fe^{2+}) → Fe^{3+}	1 × 1 = **1**

*Note: the PDF slide gives n = 1 × 2 = 2 for $Fe_2(C_2O_4)_3$ — that counts only one oxalate. With **3 oxalate ions (6 C atoms)** each changing by 1, the n-factor is **6**.*

Check ON of C in oxalate: 2x + 4(–2) = –2 ⇒ **x = +3**.

EQUIVALENT WEIGHT & NORMALITY (ADDITIONAL TOPIC)

Formulas

$$\text{Equivalent weight (E)} = \text{Molar mass (M)} / \text{n-factor}$$

$$\text{Number of gram-equivalents} = \text{mass} / \text{E} = \text{moles} \times \text{n-factor}$$

$$\text{Normality (N)} = \text{gram-equivalents} / \text{volume (L)} = \text{Molarity} \times \text{n-factor}$$

Substance	n-factor	Equivalent weight
H_2SO_4 (M = 98)	2	49
$KMnO_4$ in acidic medium (M = 158)	5	31.6
$K_2Cr_2O_7$ in acidic medium (M = 294)	6	49
$FeSO_4$ → Fe^{3+} (M = 152)	1	152

LAW OF EQUIVALENCE & TITRATIONS (ADDITIONAL TOPIC)

- In any reaction, **gram-equivalents of all reactants are equal** (and equal those of each product).

Titration

$$N_1V_1 = N_2V_2 \quad (\text{or } n_1M_1V_1 = n_2M_2V_2)$$

- **Acid–base titration:** indicator phenolphthalein / methyl orange.
- **Redox titration:** $KMnO_4$ is a **self-indicator** (pink colour at end point); $K_2Cr_2O_7$ uses diphenylamine; **iodometric/iodimetric** titrations use starch.

BALANCING REDOX EQUATIONS (ADDITIONAL TOPIC)

Oxidation-number method

1. Write skeletal equation; assign ON and find atoms whose ON changes.
2. Calculate total increase and decrease in ON; make them equal by multiplying.
3. Balance atoms other than O and H.
4. Acidic medium: balance O with H_2O and H with H^+; basic medium: add H_2O and OH^-.
5. Check charge and atom balance.

Ion-electron (half-reaction) method

- Split into oxidation and reduction half-reactions; balance each for atoms (O with H_2O, H with H^+/OH^-) and charge (with e^-); equalise electrons and add.
- Example (acidic): $MnO_4^- + 5Fe^{2+} + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$.

PRACTICE MCQS

Q1. The oxidation number of S in H_2SO_5 (Caro's acid) is

(a) +8
(b) +6
(c) +4
(d) +7

Answer: (b) — One peroxide linkage.

Q2. Oxidation state of Cr in CrO_5 is

(a) +10
(b) +8
(c) +6
(d) +5

Answer: (c)

Q3. Which can act as both oxidising and reducing agent?

(a) $KMnO_4$
(b) HNO_2
(c) CH_4
(d) HNO_3

Answer: (b) — N in +3 (intermediate).

Q4. The n-factor of $KMnO_4$ in acidic medium is

(a) 1
(b) 3
(c) 5
(d) 7

Answer: (c)

Q5. The basicity of H_3PO_3 is

(a) 1
(b) 2
(c) 3
(d) 0

Answer: (b) — One H is bonded directly to P.

Q6. The average oxidation state of C in C_3O_2 is

(a) +2
(b) +4/3
(c) 0
(d) +4

Answer: (b) — $3x - 4 = 0 \Rightarrow x = +4/3$ (atom-wise +2, 0, +2).

Q7. Oxidation number of Fe in the brown-ring complex $[Fe(H_2O)_5NO]SO_4$ is

(a) +2
(b) +3
(c) +1
(d) 0

Answer: (c) — NO is present as NO^+.

Q8. The n-factor of ferrous oxalate when oxidised to Fe^{3+} and CO_2 is

(a) 1
(b) 2
(c) 3
(d) 6

Answer: (c)

Q9. Equivalent weight of $K_2Cr_2O_7$ in acidic medium (M = 294) is

(a) 294
(b) 147
(c) 98
(d) 49

Answer: (d) — 294/6.

QUICK REVISION

- Oxidation: loss of e^-, ON ↑; reduction: gain of e^-, ON ↓.
- O = –2 (peroxide –1, superoxide –½, in OF_2/O_2F_2 +ve); F always –1.
- Coordinate bond: EN(donor) > EN(acceptor) → ignore; EN(donor) < EN(acceptor) → +2/–2.
- Peroxide traps: H_2SO_5, $H_2S_2O_8$ (S = +6), CrO_5 (+6), K_3CrO_8 (+5). Max OS in table: +8 (Os, Ru).
- Max OS → only OA; min OS → only RA; intermediate → both.
- n-factor: acid (H^+), base (OH^-), salt (total charge), redox (ΔON × atoms). $KMnO_4$: 5 (acidic), 3 (neutral), 1 (basic); $K_2Cr_2O_7$: 6.
- E = M/n; N = M × n; $N_1V_1 = N_2V_2$.

Redox is the base of electrochemistry — practise n-factors daily!