

Redox Map

Edexcel A Level Chemistry · Topic 3 Redox I

1 Oxidation numbers: the rules 3.1–3.2

- Uncombined element: 0 (Na, Cl₂, S₈)
- Monatomic ion: its charge (Mg²⁺ is +2)
- They add up to 0 in a compound, to the charge in an ion
- Always: F -1; Group 1 +1; Group 2 +2; Al +3
- H +1, but -1 in metal hydrides (NaH)
- O -2, but -1 in peroxides (H₂O₂), +2 in OF₂
- Cl -1, unless bonded to O or F

2 Roman numerals and formulae 3.9–3.10

- The numeral is the oxidation number of the element that can vary: FeCl₃ is iron(III) chloride
- Oxyanions: -ate + the central atom's oxidation number: SO₄²⁻ is sulfate(VI), NO₂⁻ nitrate(III)
- Name → formula: balance the oxidation numbers to zero: copper(I) oxide is Cu₂O
- No numeral when the oxidation number is fixed: sodium chloride

3 Oxidation and reduction 3.3–3.4, 3.11–3.12

- Oxidation:** loss of electrons; oxidation number increases
- Reduction:** gain of electrons; oxidation number decreases
- OIL RIG: Oxidation Is Loss, Reduction Is Gain
- Redox: both happen together; electrons lost = electrons gained
- Metals lose electrons to form positive ions; non-metals gain them to form negative ions

4 Oxidising and reducing agents 3.5–3.6

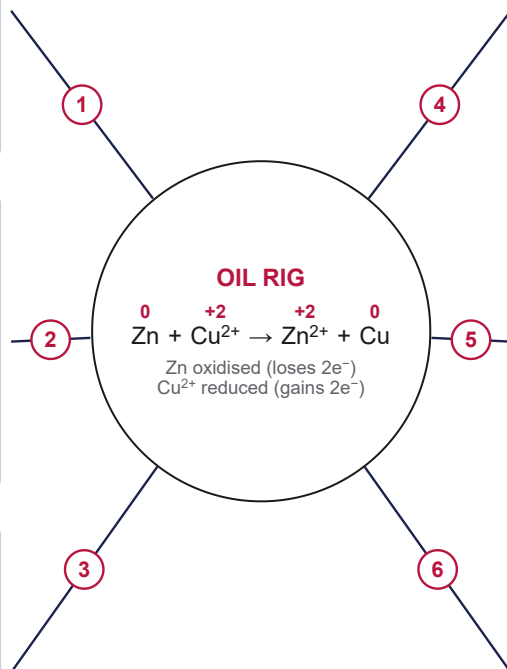
- Oxidising agent:** accepts electrons, so it is reduced (Cl₂, O₂, MnO₄⁻ in acid)
- Reducing agent:** donates electrons, so it is oxidised (metals, H₂, C, CO)
- An agent is always a reactant: name the whole species (CO, not C)

5 Half-equations 3.13

- 1 balance the element that changes
- 2 balance O with H₂O
- 3 balance H with H⁺
- 4 add e⁻ to balance the charge
- MnO₄⁻ + 8H⁺ + 5e⁻ → Mn²⁺ + 4H₂O
- Combine: make e⁻ lost = e⁻ gained, add, cancel the e⁻

6 Disproportionation 3.7–3.8

- One element in a single species is oxidised and reduced at the same time
- Cl₂ + 2NaOH → NaCl + NaClO + H₂O (cold, dilute)
- Cl: 0 → -1 in NaCl (reduced); 0 → +1 in NaClO (oxidised)
- 2Cu⁺ → Cu²⁺ + Cu
- Classify: redox if any oxidation number changes; disproportionation if one element goes both ways



Common ions: charge and oxidation number of the central atom oxidation numbers are written sign first; charges number first

Ion	Formula	Charge	Oxidation number	Systematic name
sulfate	SO ₄ ²⁻	2-	S +6	sulfate(VI)
sulfite	SO ₃ ²⁻	2-	S +4	sulfate(IV)
hydrogensulfate	HSO ₄ ⁻	1-	S +6	hydrogensulfate(VI)
thiosulfate	S ₂ O ₃ ²⁻	2-	S +2	thiosulfate
nitrate	NO ₃ ⁻	1-	N +5	nitrate(V)
nitrite	NO ₂ ⁻	1-	N +3	nitrate(III)
ammonium	NH ₄ ⁺	1+	N -3	ammonium
phosphate	PO ₄ ³⁻	3-	P +5	phosphate(V)

Ion	Formula	Charge	Oxidation number	Systematic name
carbonate	CO ₃ ²⁻	2-	C +4	carbonate
hydrogencarbonate	HCO ₃ ⁻	1-	C +4	hydrogencarbonate
hydroxide	OH ⁻	1-	O -2	hydroxide
hypochlorite	ClO ⁻	1-	Cl +1	chlorate(I)
chlorate	ClO ₃ ⁻	1-	Cl +5	chlorate(V)
permanganate	MnO ₄ ⁻	1-	Mn +7	manganate(VII)
dichromate	Cr ₂ O ₇ ²⁻	2-	Cr +6	dichromate(VI)
chromate	CrO ₄ ²⁻	2-	Cr +6	chromate(VI)

Worked examples

Working out an oxidation number

- Cr in Cr₂O₇²⁻: 2x + 7(-2) = -2, so x = +6
- S in H₂SO₄: 2(+1) + x + 4(-2) = 0, so x = +6
- Fe in Fe₃O₄: 3x + 4(-2) = 0, so x = +8/3 (a mixture of +2 and +3)

Combining: a worked example

- Cr₂O₇²⁻ + 14H⁺ + 6e⁻ → 2Cr³⁺ + 7H₂O
- Fe²⁺ → Fe³⁺ + e⁻ (× 6)
- Cr₂O₇²⁻ + 14H⁺ + 6Fe²⁺ → 2Cr³⁺ + 7H₂O + 6Fe³⁺
- Charges: +24 on each side

Write it right

- Charges: number first (2+). Oxidation numbers: sign first (+2)
- Spectator ions (Na⁺, Cl⁻, SO₄²⁻) do not change and are left out of ionic equations
- Electrons never appear in an overall equation

Oxidation numbers assume every bond is ionic: the more electronegative atom takes the shared electrons.



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