Redox Map

Cambridge International AS & A Level Chemistry · 6.1 Redox processes

1 Oxidation numbers: the rules 6.1.1

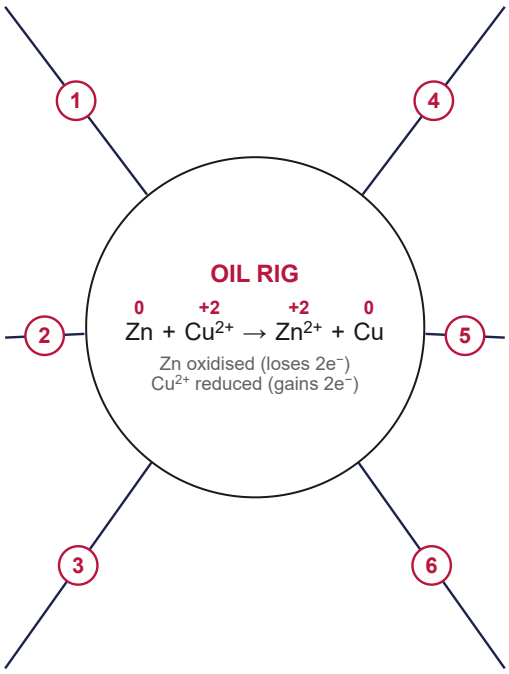
- Uncombined element: 0 (Na, Cl_2, S_8)
- Monatomic ion: its charge (Mg^{2+} is +2)
- They add up to 0 in a compound, to the charge in an ion
- Always: F −1; Group 1 +1; Group 2 +2; Al +3
- H +1, but −1 in metal hydrides (NaH)
- O −2, but −1 in peroxides (H_2O_2), +2 in OF_2
- Cl −1, unless bonded to O or F

2 Roman numerals and formulae 6.1.5

- The numeral is the oxidation number of the element that can vary: $FeCl_3$ is iron(III) chloride
- Oxyanions: -ate + the central atom's oxidation number: SO_4^{2-} is sulfate(VI), NO_2^- nitrate(III)
- Name → formula: balance the oxidation numbers to zero: copper(I) oxide is Cu_2O
- No numeral when the oxidation number is fixed: sodium chloride

3 Oxidation and reduction 6.1.3

- **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 6.1.4

- **Oxidising agent**: accepts electrons, so it is reduced (Cl_2, O_2, MnO_4^- in acid)
- **Reducing agent**: donates electrons, so it is oxidised (metals, H_2, C, CO)
- An agent is always a reactant: name the whole species (CO, not C)

5 Half-equations 6.1.2

- 1 balance the element that changes
- 2 balance O with H_2O
- 3 balance H with H^+
- 4 add e^- to balance the charge
- $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$
- Combine: make e^- lost = e^- gained, add, cancel the e^-

6 Disproportionation and balancing 6.1.3

- Disproportionation: one element in a single species is both oxidised and reduced
- $Cl_2 + 2NaOH \rightarrow NaCl + NaClO + H_2O$: Cl 0 → −1 and 0 → +1
- Balancing: the total rise in oxidation number = the total fall
- MnO_4^- (Mn +7 → +2, down 5) reacts with $5Fe^{2+}$ (each +2 → +3, up 1)

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	$S_2O_3^{2-}$	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	$Cr_2O_7^{2-}$	2−	Cr +6	dichromate(VI)
chromate	CrO_4^{2-}	2−	Cr +6	chromate(VI)

Worked examples

Working out an oxidation number

- Cr in $Cr_2O_7^{2-}$: 2x + 7(−2) = −2, so x = +6
- S in H_2SO_4: 2(+1) + x + 4(−2) = 0, so x = +6
- Fe in Fe_3O_4: 3x + 4(−2) = 0, so x = +8/3 (a mixture of +2 and +3)

Combining: a worked example

- $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$
- $Fe^{2+} \rightarrow Fe^{3+} + e^-$ (× 6)
- $Cr_2O_7^{2-} + 14H^+ + 6Fe^{2+} \rightarrow 2Cr^{3+} + 7H_2O + 6Fe^{3+}$
- Charges: +24 on each side

Write it right

- Charges: number first (2+). Oxidation numbers: sign first (+2)
- Spectator ions (Na^+, Cl^-, SO_4^{2-}) 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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