Topic 3: Redox I

Edexcel A Level Chemistry • Paper 1 (9CH0/01)

3.1 Know what is meant by the term 'oxidation number'.

Oxidation number (oxidation state) is the charge an atom would have if all bonds were ionic. It indicates the degree of oxidation of an atom.

Rules: uncombined elements = 0; monatomic ions = charge of ion; O = −2 (except peroxides = −1); H = +1 (except metal hydrides = −1); F = −1 always. Sum of oxidation numbers in a compound = 0; in an ion = charge of the ion.

3.2 Be able to calculate the oxidation number of elements in compounds and ions.

Use the rules above algebraically. E.g. in SO_4^{2-}: let S = x, O = −2. x + 4(−2) = −2, so x = +6.

Know that oxidation numbers in peroxides (O = −1, e.g. H_2O_2) and metal hydrides (H = −1, e.g. NaH) are exceptions to the usual rules.

3.3 Understand oxidation and reduction in terms of electron transfer and changes in oxidation number, applied to reactions of s- and p-block elements.

Oxidation = loss of electrons = increase in oxidation number (OIL)

Reduction = gain of electrons = decrease in oxidation number (RIG)

Remember: OIL RIG (Oxidation is Loss, Reduction is Gain)

3.4 Understand oxidation and reduction in terms of electron loss or electron gain.

In any redox reaction, oxidation and reduction occur simultaneously. The species that is oxidised loses electrons; these electrons are gained by the species that is reduced. Total electrons lost = total electrons gained.

3.5 Know that oxidising agents gain electrons.

An oxidising agent (oxidant) causes another species to be oxidised. It does this by gaining (accepting) electrons, and is therefore itself reduced in the process.

3.6 Know that reducing agents lose electrons.

A reducing agent (reductant) causes another species to be reduced. It does this by losing (donating) electrons, and is therefore itself oxidised in the process.

© Chemistry with Chloe | chemistrywithchloe.com
For personal educational use only. Not for redistribution.

3.7 Understand that a disproportionation reaction involves an element in a single species being simultaneously oxidised and reduced.

In disproportionation, the same element in one species is both oxidised and reduced simultaneously. E.g. Cl_2 + 2NaOH → NaCl + NaClO + H_2O: chlorine (0) is simultaneously reduced to −1 (in NaCl) and oxidised to +1 (in NaClO).

3.8 Know that oxidation number is a useful concept in terms of the classification of reactions as redox and as disproportionation.

If oxidation numbers change during a reaction, it is a redox reaction. If the same element in one species shows both an increase and a decrease in oxidation number, it is disproportionation. Oxidation numbers allow us to identify and classify these reactions.

3.9 Be able to indicate the oxidation number of an element in a compound or ion, using a Roman numeral.

The oxidation number is shown as a Roman numeral in brackets after the element name. E.g. iron(III) chloride = $FeCl_3$ (Fe is +3). Manganese(VII) = Mn is +7 (as in $KMnO_4$). This system is used for elements with variable oxidation states.

3.10 Be able to write formulae given oxidation numbers.

Use the oxidation numbers to work out the ratio of atoms. E.g. iron(III) oxide: Fe = +3, O = −2. To balance charges: 2 × (+3) + 3 × (−2) = 0, so Fe_2O_3. Chromium(III) sulfate: Cr = +3, SO_4 = −2, so $Cr_2(SO_4)_3$.

3.11 Understand that metals, in general, form positive ions by loss of electrons with an increase in oxidation number.

Metals are electropositive – they readily lose electrons to form cations. The oxidation number increases from 0 to a positive value. E.g. Na → Na^+ + e^- (oxidation number: 0 → +1).

3.12 Understand that non-metals, in general, form negative ions by gain of electrons with a decrease in oxidation number.

Non-metals are electronegative – they readily gain electrons to form anions. The oxidation number decreases from 0 to a negative value. E.g. Cl_2 + 2e^- → 2Cl^- (oxidation number: 0 → −1).

3.13 Be able to write ionic half-equations and use them to construct full ionic equations.

Half-equations show either the oxidation or reduction process separately.

- Balance atoms first, then balance charges using electrons.
- E.g. Mg → Mg^{2+} + 2e^- (oxidation); Cl_2 + 2e^- → 2Cl^- (reduction).
- To form the full equation, balance electrons so they cancel, then combine.
- Mg + Cl_2 → $MgCl_2$ (electrons cancel: 2 lost = 2 gained).

© Chemistry with Chloe | chemistrywithchloe.com
For personal educational use only. Not for redistribution.

ABOUT THESE NOTES

Notes written by Chloe Archard (MSc Biochemistry, Oxford-trained chemist, 20+ years teaching experience at top UK independent schools). For 1-1 tuition, revision courses, and more free resources: chemistrywithchloe.com

© Chemistry with Chloe | chemistrywithchloe.com
For personal educational use only. Not for redistribution.