Unit 6 · Oxidation and Reduction

STUDY NOTES

WHAT NEB'S GRADE 11 EXAM ASKS

- **CDC's grid (2077):** Physical Chemistry (Units 1–8 together): 28 of 75 marks
- **The paper:** 11 MCQs × 1 · 8 short × 5 · 3 long × 8; 3 hours; 1 mark per expected point
- **CDC's model paper:** short question: the Downs cell: define an electrolytic cell; mass of sodium from 0.1 A for 30 min at 75% efficiency; why calcium isn't deposited; why aqueous NaCl can't be used
- **MCQ (1 mark)**
  - what oxidation is; what is reduced; electrons moved in $2\,Fe^{3+} + Sn^{2+}$; which is true of it
  - H_2 per mol of O_2 from water (2); a disproportionation
  - charge to deposit 108 g Al (12 F); electrons for 2 mol $MnO_4^- \rightarrow Mn^{2+}$ (10)
  - overall reaction from two halves; Cu from $CuCl_2$ at 2.0 A for 30 min; KCl(aq) products; greatest mass in cells in series
- **Short (5 marks, in parts)**
  - define with an example: oxidation, reduction (electrons, oxidation number), redox, oxidizing and reducing agent
  - differences: oxidant/reductant, oxidation/reduction, electrochemical/chemical equivalent, Faraday's two laws
  - reasons: burning ethane is redox; oxidation and reduction together; an oxidant is reduced
  - oxidation numbers: NaCl, ClO_2, $Fe_2(SO_4)_3$, SO_2, I_2, $KMnO_4$, CaH_2, PO_4^{3-}; N in NF_3 … Ba_3N_2; Cl in Cl_2O, $HClO_3$, HCl, ClF_3
  - lead–acid cell and the dry cell: oxidation numbers, halves, agents
  - **balance** by oxidation number (20 equations) and ion-electron (19, acid and base)
- **Long (8 marks)**: Faraday's laws with $AuCl_3$ and $CuSO_4$ in series (current, copper's mass); silver plating (electrodes, reactions, time, electrolyte; Pt anode instead); $AgNO_3$ and $CuSO_4$ in series with 9650 C; Cu in $AgNO_3$; Cl_2 in hot NaOH (auto-redox, balance); molten NaCl at 5 A for 1 h (Na mass, Cl_2 volume, with water added, a use)

Key terms

Oxidation: loss of electrons; increase in oxidation number (classically: gain of O or an electronegative element, loss of H).

Reduction: gain of electrons; decrease in oxidation number.

Redox reaction: oxidation and reduction together; electrons lost = electrons gained.

Oxidizing agent (oxidant): takes electrons, is itself reduced.

Reducing agent (reductant): gives electrons, is itself oxidized.

Oxidation number: charge an atom would have if every bond's electrons went to the more electronegative atom.

Disproportionation (auto-redox): one element both oxidized and reduced.

Half-equation: one half of a redox change, electrons written in.

Electrolyte: compound that conducts, by moving ions, when molten or dissolved, and is decomposed by it.

Electrolysis: decomposition of an electrolyte by an electric current (a non-spontaneous change driven).

Electrolytic cell: two electrodes in an electrolyte joined to a battery; **anode** (+) oxidation, **cathode** (−) reduction.

Discharge potential: voltage at which an ion is discharged at an electrode.

Electrochemical equivalent Z: mass deposited by 1 coulomb.

Chemical (equivalent) mass E: mass deposited by 1 mol of electrons = atomic mass ÷ valency.

Faraday (F): charge of 1 mol of electrons, 96,485 ≈ 96,500 C.

Electroplating: coating a metal with another by electrolysis.

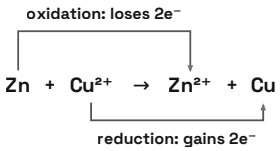
Formulas

$Q = It$	C = A × s **Use:** charge
$m = ZIt$	Faraday's first law **Use:** mass by charge
$\frac{m_1}{E_1} = \frac{m_2}{E_2}$	same charge in series **Use:** Faraday's second law
$E = \frac{A}{V} = ZF$	A atomic mass, V valency **Use:** equivalent mass
$m = \frac{EIt}{96\,500}$	for $M^{n+} + ne^-$: 1 mol M needs nF **Use:** every problem

Three ideas of oxidation

Oxidation	Reduction
gain of oxygen	loss of oxygen
gain of an electronegative element	loss of one
loss of hydrogen	gain of hydrogen
loss of electrons (OIL)	gain of electrons (RIG)
oxidation number up	oxidation number down

- $4Fe + 3O_2 \rightarrow 2Fe_2O_3$ (rust); $2Mg + O_2 \rightarrow 2MgO$
- $ZnO + H_2 \rightarrow Zn + H_2O$; $2CuO + C \rightarrow 2Cu + CO_2$
- $Fe + S \rightarrow FeS$; $2KI + Cl_2 \rightarrow 2KCl + I_2$; $2FeCl_3 + H_2 \rightarrow 2FeCl_2 + 2HCl$



Zinc gives two electrons to each copper ion: zinc oxidized, Cu^{2+} reduced.

- $Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$: halves $Zn \rightarrow Zn^{2+} + 2e^-$ and $Cu^{2+} + 2e^- \rightarrow Cu$; SO_4^{2-} a spectator; blue fades
- $2Na + Cl_2 \rightarrow 2NaCl$: **$Cl_2$ oxidizes Na** (oxidant); Na is the reductant
- burning ethane: C −3 → +4 (oxidized), O 0 → −2 (reduced)
- dry cell: $Zn \rightarrow Zn^{2+} + 2e^-$ (case); $2MnO_2 + 2NH_4^+ + 2e^- \rightarrow Mn_2O_3 + 2NH_3 + H_2O$ (Mn +4 → +3)
- oxidants: O_2, Cl_2, F_2, $KMnO_4$, $K_2Cr_2O_7$, HNO_3, H_2O_2; reductants: Na, Mg, Zn, H_2, C, CO, H_2S
- strength by E°: F_2 (+2.87 V) strongest oxidant; Na (−2.71) strong reductant; a left-side species reacts with a right-side one lower down

Oxidation numbers

1. element alone **0** (Cl_2, P_4, S_8, O_3, Fe)
2. simple ion = its charge
3. F always −1; group 1 +1; group 2 +2; Al +3
4. H +1 with non-metals; **−1 in metal hydrides** (NaH, CaH_2)
5. O −2; **peroxides −1** (H_2O_2, Na_2O_2, BaO_2); superoxides −½ (KO_2); OF_2 +2
6. sum = 0 for a compound, = charge for an ion

- $KMnO_4$: 1 + x − 8 = 0 → Mn **+7**; $Cr_2O_7^{2-}$: 2x − 14 = −2 → **+6**
- CO_3^{2-} C +4; $C_6H_{12}O_6$ C 0; $Na_2S_2O_3$ S +2
- ClO_2 Cl **+4**; ICl_3 I +3; Fe_3O_4 +8/3 (average); $S_4O_6^{2-}$ +2.5
- Cl_2O_4 (Cl–O–ClO_3): +7 and +1; $H_2S_2O_5$: +5 and +3
- N: NF_3 +3, N_2O +1, N_2H_4 −2, NO +2, N_2O_4 +4, NO_2^- +3, Ba_3N_2 −3
- Cl: Cl_2O +1, $HClO_3$ +5, HCl −1, ClF_3 +3; higher number, stronger oxidant ($HClO_3$)
- range: group 14 −4 to +4, 15 −3 to +5, 16 −2 to +6, 17 −1 to +7
- Stock names: iron(III) chloride, copper(I) oxide, sulphuric(VI) acid, potassium manganate(VII)

Disproportionation

- $Cl_2 + 2\,NaOH \longrightarrow NaCl + NaClO + H_2O$ (cold): Cl 0 → −1 and +1
- $3\,Cl_2 + 6\,NaOH \longrightarrow 5\,NaCl + NaClO_3 + 3\,H_2O$ (hot, conc.): 0 → −1 and +5
- $2\,H_2O_2 \longrightarrow 2\,H_2O + O_2$: O −1 → −2 and 0

Balancing

Oxidation number method

1. oxidation numbers; find the atoms that change
2. total change per formula ($Cr_2O_7^{2-}$: 2 × 3 = 6)
3. cross-multiply so the increase = the decrease
4. balance the changed atoms, then others: metals, non-metals, H, O last (H^+/OH^- and H_2O for ions)

- $3CuO + 2NH_3 \rightarrow 3Cu + N_2 + 3H_2O$ (N up 3 × 2, Cu down 2 × 3)
- $S + 6HNO_3 \rightarrow H_2SO_4 + 6NO_2 + 2H_2O$

Ion-electron method

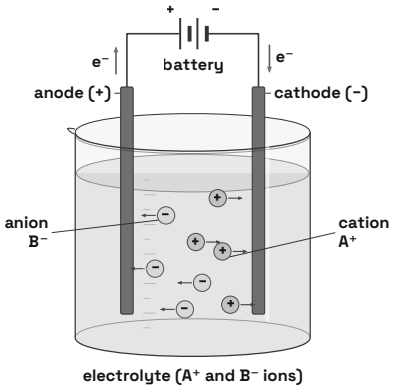
1. split into oxidation and reduction halves
2. balance atoms other than O, H
3. O with H_2O; H with H^+
4. charge with e^-
5. equal electrons; add; cancel
6. in **base**: add OH^- to both sides for each H^+, make H_2O, cancel

- $MnO_4^- + 8\,H^+ + 5\,e^- \longrightarrow Mn^{2+} + 4\,H_2O$
- $Cr_2O_7^{2-} + 14\,H^+ + 6\,e^- \longrightarrow 2\,Cr^{3+} + 7\,H_2O$
- $MnO_4^- + 8\,H^+ + 5\,Fe^{2+} \longrightarrow Mn^{2+} + 5\,Fe^{3+} + 4\,H_2O$: charge +17 both sides

- oxalate: $2\,MnO_4^- + 5\,C_2O_4^{2-} + 16\,H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$
- check atoms and charge; in base no H^+ may remain

Electrolysis

- strong electrolytes: strong acids, bases, salts; weak: CH_3COOH, $NH_3(aq)$; non: sugar, glucose, urea, ethanol
- ions exist before the current; it only moves them; solids don't conduct (ions fixed)
- **An Ox, Red Cat**: anode oxidation, reduction cathode; cations → cathode, anions → anode



Cations gain electrons at the cathode (−); anions lose them at the anode (+).

What forms

- molten NaCl: Na at the cathode, Cl_2 at the anode (the Downs cell: + $CaCl_2$ to melt at ~600 °C, graphite anode, steel cathode)
- aqueous: H^+ (water) beats $Na^+ \rightarrow H_2$; concentrated Cl^- beats $OH^- \rightarrow Cl_2$; NaOH left (chlor-alkali); very dilute → O_2
- cations hardest → easiest: Li^+ K^+ Ca^{2+} Na^+ Mg^{2+} Al^{3+} Zn^{2+} Fe^{2+} Ni^{2+} **H^+** Cu^{2+} Hg^{2+} Ag^+ Au^{3+}
- anions: $SO_4^{2-} > NO_3^- > OH^- > Cl^- > Br^- > I^-$
- Downs cell: Ca^{2+} harder to discharge than Na^+, so sodium forms; with water, H_2 would form instead of Na
- water (+ dil. H_2SO_4): $H_2 : O_2 = 2 : 1$

Electrolyte, electrodes	Anode (+)	Cathode (−)
NaCl molten, inert	Cl_2	Na
NaCl conc. aq., inert	Cl_2	H_2
NaOH aq., inert	O_2	H_2
$CuSO_4$ aq., inert	O_2 (solution acidic, blue fades)	Cu
$CuSO_4$ aq., copper	Cu dissolves (blue stays)	Cu
$AgNO_3$ aq., silver	Ag dissolves	Ag

- copper electrodes: refining (impure anode, pure cathode, anode mud); plating: object = cathode, plating metal = anode, its salt solution (silver: $AgNO_3$; amount of $AgNO_3$ unchanged)
- uses: Na, Mg, Al from molten compounds; Cl_2, NaOH, H_2 from brine; plating; refining; anodizing; hydrogen fuel

Faraday's laws

- **first**: mass ∝ charge, $m = ZIt$; Cu: $Z = 3.29 \times 10^{-4}$ g/C
- **second**: same charge in series → masses ∝ equivalent masses
- 1 F gives 1 g H_2, 31.75 g Cu, 108 g Ag, 9 g Al
- method: Q → mol e^- (÷ F) → mol product (half-equation) → mass or volume
- 108 g Al = 4 mol, needs 12 F; 0.1 mol Ag in series → 0.05 mol Cu
- 9.85 g Au (Au^{3+}, E = 65.67) in 5 h → 0.804 A; Cu in series 4.76 g
- 5.0 A, 2 h: 0.373 mol e^- → 0.187 mol H_2, 2.09 L O_2 at STP
- efficiency: actual = theoretical × %

COMMON MISTAKES

✗ $4Fe + 3O_2 \rightarrow Fe_2O_3$
✓ $4Fe + 3O_2 \rightarrow \mathbf{2}Fe_2O_3$; balance every equation.

✗ $ZnO + H_2$ gives zinc and oxygen
✓ zinc and **water**.

✗ sodium oxidizes chlorine
✓ chlorine oxidizes sodium; Na is the reducing agent.

✗ group 1 metals 0 in compounds
✓ +1 (0 only as the element).

✗ Cl in ClO_2 is +7
✓ +4.

✗ most hydrogen comes from electrolysis
✓ from natural gas.

✗ the current dissociates the electrolyte
✓ the ions are there already; the current moves them.

✗ the Downs cell uses Pt electrodes
✓ graphite anode, steel cathode.

✗ aluminium deposited from its solution
✓ only from molten Al_2O_3 (water gives H_2).