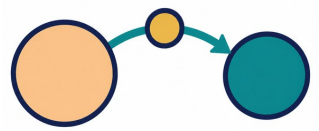
Electrode Potentials Map

AQA A-level Chemistry · 3.1.11 Electrode potentials and electrochemical cells

1 Redox and half-equations 3.1.11.1

- Oxidation: **loss** of e^-, ox. state up
- Balance O with H_2O, H with H^+
- Balance charge with e^-
- Match electrons, then add



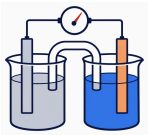
2 Standard electrode potentials 3.1.11.1

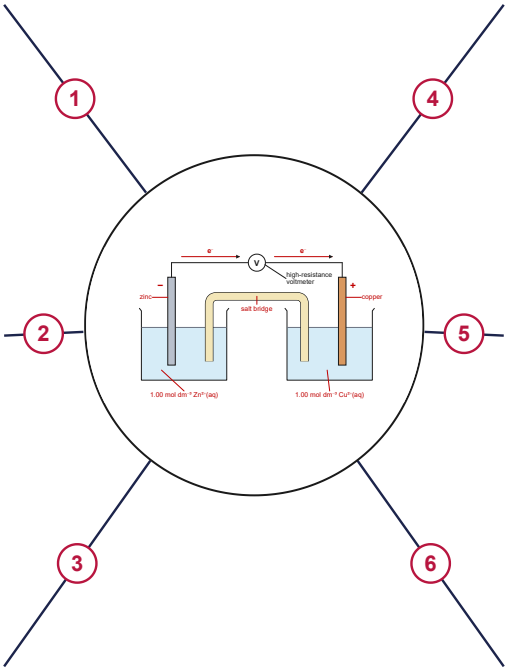
- SHE: H_2 100 kPa, 1.00 mol dm^{-3} H^+, Pt
- **298 K**; E⊖(SHE) = 0.00 V
- E⊖: emf against the SHE
- Written as **reductions**



3 Cells and E⊖cell 3.1.11.1

- Salt bridge KNO_3: **ions** move
- E⊖cell = E⊖(right) − E⊖(left)
- Zn | Zn^{2+} ‖ Cu^{2+} | Cu
- Pt for gases and ion pairs





4 Series and feasibility 3.1.11.1

- Most positive: best **oxidising** agent
- Most negative: best **reducing** agent
- E⊖cell **positive**: feasible
- Choose a reagent: the E⊖ **window**



5 Commercial cells 3.1.11.2

- More negative couple runs **backwards**
- EMF = E(+) − E(−) from data
- Recharging **reverses** both reactions
- Lithium cell: Li → $Li^+ + e^-$



6 Fuel cells, benefits and risks 3.1.11.2

- $H_2 + O_2$ supplied **continuously**
- (−) $H_2 + 2OH^- \rightarrow 2H_2O + 2e^-$
- (+) $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$
- **Risks**: flammable, toxic metals



Some standard electrode potentials

data booklet values, 298 K

Half-equation	E⊖ / V	Strongest
$F_2 + 2e^- \rightleftharpoons 2F^-$	+2.87	oxidising agent
$Cl_2 + 2e^- \rightleftharpoons 2Cl^-$	+1.36	
$Fe^{3+} + e^- \rightleftharpoons Fe^{2+}$	+0.77	
$Cu^{2+} + 2e^- \rightleftharpoons Cu$	+0.34	
$2H^+ + 2e^- \rightleftharpoons H_2$	0.00	reference
$Zn^{2+} + 2e^- \rightleftharpoons Zn$	−0.76	reducing agent

Worked examples

Cell potential

- Zn | Zn^{2+} ‖ Cu^{2+} | Cu
- +0.34 − (−0.76)
- **+1.10 V**

Feasible?

- $Cl_2 + 2Fe^{2+}$
- 1.36 − 0.77 = +0.59 V
- **feasible**

Half-equation

- $MnO_4^- \rightarrow Mn^{2+}$
- $+ 8H^+ + 5e^-$, $+ 4H_2O$
- **balanced**

Redox titrations with manganate(VII) follow in 3.2.5.5; transition-metal redox returns in 3.2.5.

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