Elements from the Sea (ES)

Revision guide

Clarity, confidence, and top grades.

What this storyline covers

- Atom economy
- Oxidation states and redox
- Electrolysis and electrode half-equations
- The halogens: properties and reactivity
- Halide displacement reactions
- Silver nitrate tests for halides
- Hydrogen halides and their reactions
- Dynamic equilibrium and K_c

Required practicals in this storyline: **PAG4** (halide tests), iodine–thiosulfate titrations, electrolysis of aqueous solutions, and redox test-tube reactions.



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How to use this guide

Elements from the Sea brings together **redox**, **electrolysis** and **equilibrium** — three ideas that reappear throughout A-level — using the halogens as the context. The redox and equilibrium work here is built on again in The Chemical Industry, Developing Metals and Oceans, so it is worth securing now.

How it is assessed

OCR B is context-led, so ES content can appear in **any** of the three written papers, in any context:

- **Paper 1** — Fundamentals of chemistry (110 marks)
- **Paper 2** — Scientific literacy, including the pre-release Advance Notice article (100 marks)
- **Paper 3** — Practical skills in chemistry (60 marks)

How to revise it

- Learn the **definitions in the teal boxes** word-for-word — they are marks you can guarantee.
- Work every **worked example** yourself before reading the solution.
- Learn the **colours and observations** exactly — halogen, precipitate and end-point colours are common marks.
- Use the **spec checklist** on the last page to confirm you can do every point.

1 Atom economy

Spec reference: ES(a).

Key definition

Atom economy = (M_r of the desired product ÷ sum of the M_r of all products) × 100%. It measures how much of the reactant mass ends up as the useful product.

- A **high** atom economy means less waste and a more **sustainable, efficient** process.
- **Addition** reactions have 100% atom economy (a single product); substitution and elimination reactions are lower.

Worked example — atom economy

Find the atom economy for ethanol from fermentation: $C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$ (desired product ethanol).

M_r: ethanol = 46.0 → 2 × 46.0 = 92.0; CO_2 = 44.0 → 2 × 44.0 = 88.0

atom economy = 92.0 / (92.0 + 88.0) × 100 = 92.0 / 180.0 × 100 = 51.1%

2 Oxidation states and redox

Spec references: ES(d), ES(e), ES(f), ES(g). Practical: iodine–thiosulfate titration.

Oxidation, reduction and agents ES(d)

Key definitions

- **Oxidation** is **loss** of electrons; **reduction** is **gain** of electrons (OIL RIG).
- An **oxidising agent** accepts electrons (and is itself reduced); a **reducing agent** donates electrons (and is itself oxidised).

Write **half-equations** for the oxidation and the reduction, balancing the atoms and then adding electrons; combine them so the electrons cancel to give the overall equation.

Oxidation states ES(e)

Rule	Oxidation state
Uncombined element	0
Simple monatomic ion	the ionic charge
Group 1 / Group 2 / aluminium	+1 / +2 / +3
Fluorine	always −1
Hydrogen	+1 (−1 in metal hydrides)
Oxygen	−2 (−1 in peroxides)

- The oxidation states in a formula add up to **0** (neutral compound) or to the **overall charge** (ion).
- A species is **oxidised** if its oxidation state **increases**, and **reduced** if it **decreases**.

Worked example — identifying redox

In $Cl_2 + 2NaBr \rightarrow 2NaCl + Br_2$, what is oxidised and what is reduced?

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Cl: 0 → −1 (gains electrons) → chlorine is reduced; it is the oxidising agent
Br: −1 → 0 (loses electrons) → bromide is oxidised; it is the reducing agent
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Balancing redox by oxidation states ES(f)

- For redox that does **not** involve acid–base, balance so that the **total increase** in oxidation state equals the **total decrease** (electrons lost = electrons gained).
- Example: $3Ca + 2Al^{3+} \rightarrow 3Ca^{2+} + 2Al$ (each Ca loses $2e^-$, each Al^{3+} gains $3e^-$; $3 \times 2 = 2 \times 3 = 6$ electrons transferred).

Iodine–thiosulfate titration ES(f)

- Iodine reacts with thiosulfate: $I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$.
- Add **starch** indicator near the end point; the blue-black colour disappears (turns colourless) at the end point.
- Used to find the concentration of an **oxidising agent**, which first releases iodine from excess iodide; the iodine is then titrated.

Worked example — iodine–thiosulfate titration

$25.0\ cm^3$ of an iodine solution needed $22.0\ cm^3$ of $0.100\ mol\ dm^{-3}$ sodium thiosulfate. Find $[I_2]$.

$n(S_2O_3^{2-}) = 0.100 \times 22.0/1000 = 2.20 \times 10^{-3}$ mol

$n(I_2) = \frac{1}{2} \times 2.20 \times 10^{-3} = 1.10 \times 10^{-3}$ mol (ratio 1 : 2)

$[I_2] = (1.10 \times 10^{-3}) / (25.0/1000) = 0.0440\ mol\ dm^{-3}$

Naming inorganic compounds ES(g)

- Roman numerals show the **oxidation state** of the named element: copper(II) sulfate, iron(III) chloride.
- In an **-ate** ion the numeral gives the oxidation state of the central atom: sodium chlorate(I) NaClO, lead(II) nitrate(V) $Pb(NO_3)_2$, potassium manganate(VII) $KMnO_4$.

3 Electrolysis and extracting the halogens

Spec references: ES(b), ES(c). Practical: electrolysis of aqueous solutions.

Extracting the halogens from the sea ES(b)

- **Chlorine** is made by the **electrolysis of brine** (concentrated aqueous sodium chloride).
- **Bromine** is obtained from seawater by **displacement** with chlorine: $Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$.
- You only need to **interpret given information** about these processes — recall of the industrial detail is not required.

Electrolysis ES(c)

Key idea

In **electrolysis**, a direct current decomposes a molten or aqueous ionic compound. **Reduction** (gain of electrons) happens at the **cathode (−)**; **oxidation** (loss of electrons) happens at the **anode (+)**.

Molten ionic compounds

- The metal forms at the cathode and the non-metal at the anode. Molten NaCl: cathode $Na^+ + e^- \rightarrow Na$; anode $2Cl^- \rightarrow Cl_2 + 2e^-$.

Aqueous solutions

Water also supplies H^+ and OH^-, so the products follow these rules:

Electrode	Rule	Half-equation
Cathode (−)	Group 1, Group 2 and aluminium salts give **hydrogen**; any other metal is **plated**	$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ or $M^{n+} + ne^- \rightarrow M$
Anode (+)	a **halide** gives the **halogen**; otherwise **oxygen** forms	$2Cl^- \rightarrow Cl_2 + 2e^-$ or $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$

Worked example — electrode products

Predict the products of electrolysing (a) aqueous NaCl and (b) aqueous $CuSO_4$.

(a) NaCl(aq): Na is Group 1 → cathode $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$; anode $2Cl^- \rightarrow Cl_2 + 2e^-$

(b) $CuSO_4$(aq): Cu is not Gp 1/2/Al → cathode $Cu^{2+} + 2e^- \rightarrow Cu$; anode (sulfate) $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$

4 The halogens

Spec references: ES(h), ES(i), ES(j).

Physical properties ES(h)

Halogen	Appearance at room temperature
Fluorine, F_2	pale yellow gas
Chlorine, Cl_2	pale green gas
Bromine, Br_2	red-brown liquid
Iodine, I_2	grey-black solid (purple vapour)

- **Volatility decreases** down the group (boiling points rise) as the molecules get larger and the induced-dipole–induced-dipole bonds get stronger.
- The halogens are only **slightly soluble in water** but dissolve more readily in **organic solvents**, giving characteristic colours (chlorine pale green, bromine orange, iodine violet).

Reactivity ES(i)

- The halogens are **oxidising agents**: each gains one electron to form a 1– ion.
- **Reactivity decreases down the group**: the atom gets larger with more shielding, so the incoming electron is attracted **less strongly** and gained **less easily**.

Displacement reactions with halide ions ES(j)

- A **more reactive** halogen **displaces** a less reactive halide from solution (it is the stronger oxidising agent).
- $Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$ (turns orange); $Cl_2 + 2I^- \rightarrow 2Cl^- + I_2$ (brown); $Br_2 + 2I^- \rightarrow 2Br^- + I_2$ (brown). Iodine displaces neither chloride nor bromide.
- Half-equations (Cl_2 with Br^-): $Cl_2 + 2e^- \rightarrow 2Cl^-$ (reduction); $2Br^- \rightarrow Br_2 + 2e^-$ (oxidation).
- The colours are clearer in an **organic solvent** layer added after shaking.

5 Testing for halide ions

Spec reference: ES(k). Practical: PAG4.

Add dilute **nitric acid**, then **silver nitrate** solution, to the halide in solution. A silver halide precipitate forms: $Ag^+(aq) + X^-(aq) \rightarrow AgX(s)$.

Halide	Precipitate with $AgNO_3$	Solubility in ammonia
Chloride, Cl^-	white (AgCl)	dissolves in dilute ammonia
Bromide, Br^-	cream (AgBr)	dissolves in concentrated ammonia
Iodide, I^-	yellow (AgI)	insoluble even in concentrated ammonia

> **Exam tip**
> The acid is added first to remove carbonate or hydroxide ions, which would otherwise give their own precipitates. The **ammonia step** lets you tell the three precipitates apart when their colours are hard to judge.

6 Hydrogen halides and chlorine

Spec references: ES(l), ES(m), ES(n).

Preparing the hydrogen halides ES(l)

- **HCl** is made from a chloride and **concentrated sulfuric acid**: $NaCl + H_2SO_4 \rightarrow NaHSO_4 + HCl$.
- **HBr and HI cannot** be made this way, because sulfuric acid **oxidises** them. Instead use the halide salt with **phosphoric acid**, a non-oxidising acid, e.g. $NaBr + H_3PO_4 \rightarrow NaH_2PO_4 + HBr$ (the phosphoric acid equations are not required).

Action of sulfuric acid on solid halides ES(m)

Halide	What happens with concentrated H_2SO_4
Chloride	only HCl produced (acid–base, no redox)
Bromide	HBr produced, then some reduces H_2SO_4 to **SO_2**; brown Br_2 also forms
Iodide	HI produced, then reduces H_2SO_4 further to **H_2S** (and S); purple I_2 also forms

This shows the **reducing power increases** down the group: HI is the strongest reducing agent.

Properties of the hydrogen halides ES(m)

- **Thermal stability decreases** down the group (the H–X bond weakens), so HCl is the most stable and HI the least.
- They **all** react with **ammonia** to give a white solid (HX + NH_3 → NH_4X) and dissolve in water to give **acids**.
- They differ in their **reactions with sulfuric acid**, which reflects their reducing power (above).

Uses and risks of chlorine ES(n)

- Chlorine is **toxic**, so its storage and transport carry real risks.
- These are weighed against its uses: **sterilising water** by killing bacteria (a major public-health benefit) and **bleaching**. For drinking water, the benefit is generally judged to outweigh the risk.

7 Dynamic equilibrium and K_c

Spec references: ES(o), ES(p), ES(q). Practical: qualitative equilibrium experiments.

Dynamic equilibrium ES(o)

Key definition

At **dynamic equilibrium** in a closed system, the **forward and reverse reactions occur at the same rate**, so the concentrations of reactants and products stay constant while both reactions continue.

The equilibrium constant K_c ES(p)

$$\text{for } aA + bB \rightleftharpoons cC + dD, \quad K_c = [C]^c[D]^d \div [A]^a[B]^b$$

equilibrium concentrations in mol dm^{-3}; units are not required

- A **large** K_c (>> 1) means the equilibrium lies well to the **right** (mostly products); a **small** K_c (<< 1) means it lies to the **left**.

Worked example — calculating K_c

For $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ at equilibrium in a 1.0 dm^3 flask: $[H_2] = 0.20$, $[I_2] = 0.20$, $[HI] = 1.60$ mol dm^{-3}.

$K_c = [HI]^2 / ([H_2][I_2]) = (1.60)^2 / (0.20 \times 0.20)$
$= 2.56 / 0.040 = 64$

Changing the conditions ES(q)

- **Concentration:** K_c is fixed at constant temperature, so changing a concentration shifts the position to restore the ratio. If a term on the top is increased, a term on the bottom must increase too, so the position moves to the **left**.
- **Temperature:** changes the **value** of K_c. For an exothermic forward reaction, raising the temperature **decreases** K_c (position moves left); for an endothermic reaction it increases K_c.
- **Pressure** (gases): shifts the position towards the side with **fewer gas molecules**, but does **not** change K_c.
- In each case the system opposes the change imposed on it.

P Required practicals in this storyline

These PAG activities are most likely to be examined in Paper 3, but the techniques can appear in any paper.

Activity	What to be able to do
Halide tests (PAG4)	add nitric acid then silver nitrate; identify Cl^- (white), Br^- (cream), I^- (yellow); confirm with dilute and concentrated ammonia
Iodine–thiosulfate titration	titrate iodine against thiosulfate, adding starch near the end point; calculate the concentration of an oxidising agent
Electrolysis of aqueous solutions	predict and observe the products at each electrode and write the half-equations
Redox / displacement reactions	carry out test-tube halogen and halide reactions; record colours and deduce the reactivity order

Exam tip — practical accuracy language

Precise **observations** and **state symbols** earn marks: describe colours (including in the organic layer), add the **starch indicator near the end point** of a thiosulfate titration, and identify the apparatus with the largest **percentage uncertainty**.

Spec coverage checklist — can you do every point?

Every assessable learning outcome for Elements from the Sea, with where it is covered in this guide. Tick each one once you can do it confidently.

ES(a)	Atom economy and its relation to the efficient use of atoms — §1
ES(b)	Interpret given information on extracting the halogens from sea minerals — §3
ES(c)	Electrolysis of molten and aqueous salts; electrode half-equations — §3
ES(d)	Redox as electron transfer; half-equations; oxidising and reducing agents — §2
ES(e)	Assign and calculate oxidation states; identify what is oxidised and reduced — §2
ES(f)	Balance redox equations by oxidation states; iodine–thiosulfate titrations — §2
ES(g)	Systematic nomenclature of inorganic compounds — §2
ES(h)	Physical properties of the halogens — §4
ES(i)	Relative reactivity of the halogens as electron gain — §4
ES(j)	Redox displacement reactions of halogens with halide ions — §4
ES(k)	Halide reactions with silver ions; precipitate colours and ammonia solubility — §5
ES(l)	Preparation of HCl, HBr and HI; action of sulfuric acid on halides — §6
ES(m)	Properties of the hydrogen halides: stability, ammonia, acidity, sulfuric acid — §6
ES(n)	Risks and uses of chlorine — §6
ES(o)	Characteristics of dynamic equilibrium — §7
ES(p)	The equilibrium constant K_c; calculations and the meaning of its magnitude — §7
ES(q)	Use K_c to explain the effects of concentration, temperature and pressure — §7