

ESAT

Chemistry

Redox and electrolysis

Free lesson notes

Covers	C5.1 to C5.6 and C12.1 to C12.6
Format	Lesson, worked examples, self-check
Price	Free

About this lesson

Half of chemistry is about electrons being moved from one place to another. When magnesium burns, when iron rusts, when a metal is extracted from its ore, when a battery runs or a spoon is silver-plated, the same thing is happening: electrons are transferred. Chemists give this one idea two names. When a reaction *happens by itself* and transfers electrons we call it **redox**. When we *force* electrons to move with a battery, to split up a compound, we call it **electrolysis**.

We start with the oldest definition of oxidation and reduction (gaining and losing oxygen), upgrade it to the electron definition, and then add **oxidation states**, a bookkeeping system that lets you spot electron transfer in any equation, even where no ions are obvious. After that comes electrolysis: what the apparatus is, why the current must be direct, how to predict the products for molten and aqueous electrolytes, how to write the half-equations, and two applications, electroplating and the extraction of aluminium.

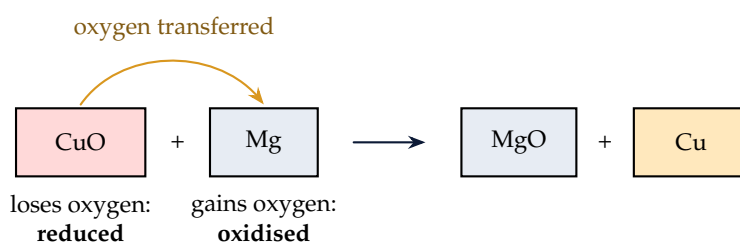
How to use it. Cover each worked example and try it before reading the solution. Everything in this lesson can be done without a calculator. Where a reactivity series is needed, we use K, Na, Ca, Mg, Al, (C), Zn, Fe, (H), Cu, Ag, Au, from most to least reactive.

Heurisco is an independent practice provider and is not affiliated with, endorsed by, or connected to UAT-UK, Pearson VUE, the UCAT Consortium, or any university. All questions are original.

1 Oxidation and reduction as oxygen transfer (C5.1)

The first definitions come from the oldest kind of chemistry, burning and smelting. **Oxidation** is the *gain of oxygen* by a substance; **reduction** is the *loss of oxygen*. When magnesium burns in air, $2\text{Mg} + \text{O}_2 \longrightarrow 2\text{MgO}$, the magnesium gains oxygen so it is oxidised. When copper(II) oxide is heated in a stream of hydrogen, $\text{CuO} + \text{H}_2 \longrightarrow \text{Cu} + \text{H}_2\text{O}$, the copper oxide loses oxygen (it is reduced) and the hydrogen gains it (it is oxidised).

The key insight is that the two always happen **together**. Oxygen cannot just disappear from one substance; it must go somewhere. Whenever one substance is reduced, another is oxidised, and a reaction in which both happen is called a **redox** reaction (*reduction and oxidation*).



Worked example: tracking oxygen

In the reaction $\text{CuO} + \text{Mg} \longrightarrow \text{MgO} + \text{Cu}$, which substance is oxidised and which is reduced? If 8.0 g of copper(II) oxide is completely reduced, what mass of copper is formed? (A_r : O = 16, Cu = 64.)

Copper oxide loses oxygen so it is reduced; magnesium gains oxygen so it is oxidised. $M_r(\text{CuO}) = 80$, so $8.0/80 = 0.10$ mol of CuO, giving 0.10 mol of Cu: mass = $0.10 \times 64 = 6.4$ g. (The missing 1.6 g is the oxygen that moved to the magnesium.)

Common mistakes

- Thinking oxidation always involves oxygen. This definition is only the first of three; it breaks down in reactions with no oxygen at all, such as $\text{Mg} + \text{Cl}_2 \longrightarrow \text{MgCl}_2$ (next section).
- Calling a reaction “an oxidation” without noticing that something else is being reduced. Redox is always a two-sided process.
- Mixing up which substance is which: the oxide that *loses* oxygen is reduced, but it is the metal or hydrogen that *gains* it that does the reducing.

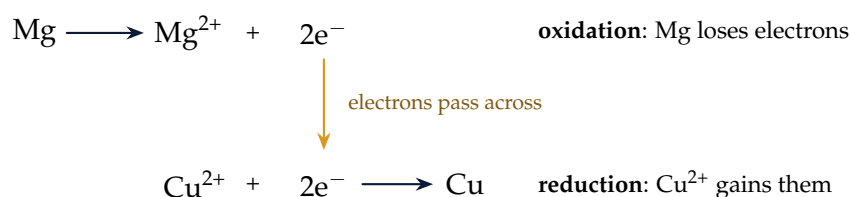
2 Oxidation and reduction as electron transfer (C5.2)

Burning magnesium in chlorine gives MgCl_2 , which is clearly the same kind of reaction as burning it in oxygen, yet there is no oxygen. What the two have in common is that magnesium atoms lose two electrons each to become Mg^{2+} ions. So we redefine the terms in a way that works for every case:

Oxidation is loss of electrons. Reduction is gain of electrons.

The memory aid is **OIL RIG**: Oxidation Is Loss, Reduction Is Gain (of electrons). Check this against section 1: when magnesium reacts with oxygen, $\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$, while oxygen atoms take two electrons each: $\text{O} + 2\text{e}^- \rightarrow \text{O}^{2-}$. Gaining oxygen and losing electrons are the same event seen from two sides.

Half-equations. Because electrons leave one species and arrive at another, we split a redox equation into two **half-equations**, one for each. Take magnesium added to copper(II) sulfate solution, a *displacement* reaction: $\text{Mg} + \text{Cu}^{2+} \rightarrow \text{Mg}^{2+} + \text{Cu}$ (the sulfate ions are unchanged *spectator ions* and are left out). The two halves are



Add the half-equations and the electrons *cancel*: they must, because none is left over. A half-equation must balance in **atoms and in charge**. In $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$ the left has charge $+2 - 2 = 0$ and so does the right. When the two halves involve different numbers of electrons, multiply one (or both) so that the numbers match before adding.

Worked example: combining half-equations

Aluminium reacts with copper(II) ions. The half-equations are $\text{Al} \rightarrow \text{Al}^{3+} + 3\text{e}^-$ and $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$. Write the overall ionic equation.

Aluminium gives 3 electrons, copper takes 2; the lowest common multiple is 6. Multiply the first by 2 and the second by 3: $2\text{Al} \rightarrow 2\text{Al}^{3+} + 6\text{e}^-$ and $3\text{Cu}^{2+} + 6\text{e}^- \rightarrow 3\text{Cu}$. Add: $2\text{Al} + 3\text{Cu}^{2+} \rightarrow 2\text{Al}^{3+} + 3\text{Cu}$. Check the charge: left $3 \times (+2) = +6$; right $2 \times (+3) = +6$.

Common mistakes

- Putting the electrons on the wrong side: electrons appear on the *right* of an oxidation half-equation and on the *left* of a reduction half-equation.
- Balancing atoms but not charge. $\text{Cu}^{2+} \rightarrow \text{Cu} + 2\text{e}^-$ looks plausible; check that the charges match ($+2$ against $0 - 2 = -2$ shows it is wrong).
- Leaving electrons in the overall equation. If any remain, the two halves have not been multiplied to the same number of electrons.

3 Oxidation states (C5.3, C5.4)

Electron transfer is easy to see when ions form, but what about $\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}$, where the bonds are covalent and no electron actually leaves its molecule? We use a bookkeeping device: pretend every bond is ionic, giving the shared electrons to the more electronegative atom. The charge each atom would then have is its **oxidation state** (also called oxidation number). It is written with the sign first: $+2$, -1 , 0 . (In names, the Roman numeral gives the same number: iron(III) is Fe^{3+} , oxidation state $+3$.) The rules, applied from the top down, are:

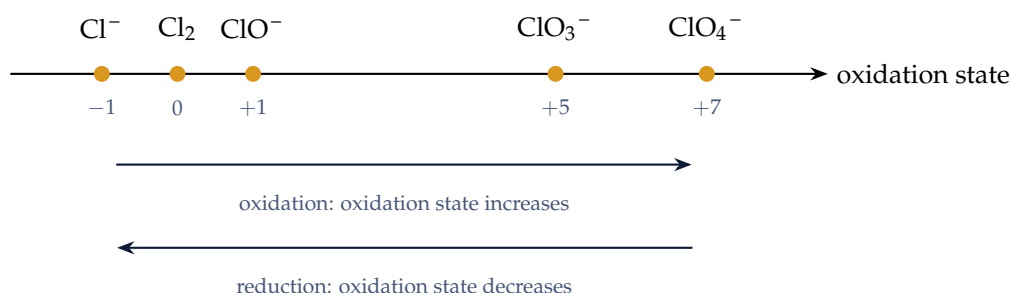
- 1 An atom in an uncombined element has oxidation state 0 (Fe, O₂, Cl₂, S₈).
- 2 A simple ion has an oxidation state equal to its charge (Na⁺ is +1, S²⁻ is -2).
- 3 The oxidation states in a compound add up to 0; in an ion they add up to the charge on the ion.
- 4 Group 1 metals are +1, group 2 metals +2, aluminium +3, fluorine -1.
- 5 Hydrogen is +1 (but -1 in metal hydrides such as NaH).
- 6 Oxygen is -2 (but -1 in peroxides such as H₂O₂).
- 7 Chlorine, bromine and iodine are -1 in compounds, except when combined with oxygen or fluorine.

To find an unknown oxidation state, write x for it, add up all the atoms, and set the total equal to the charge. In sulfuric acid, H₂SO₄: $2(+1) + x + 4(-2) = 0$, so $x = +6$. The Roman numeral appears in names such as sulfur(VI) oxide for this reason.

Using oxidation states to spot redox (C5.4). Compare the oxidation state of each element before and after the reaction:

- an element whose oxidation state **increases** is **oxidised** (it has lost electrons or, in covalent bonding, has effectively lost them);
- an element whose oxidation state **decreases** is **reduced**;
- if *nothing* changes oxidation state, the reaction is **not** redox: neutralisation, precipitation and thermal decomposition of carbonates are examples.

The change always balances: the total increase equals the total decrease, because electrons are conserved.



The picture above is a “ladder” for chlorine. Going *up* the ladder (more positive) is oxidation; going *down* is reduction. It shows why one element can be oxidised or reduced depending on what it reacts with: chlorine gas at 0 can be reduced to chloride (-1) or oxidised to chlorate(V) (+5).

Worked example: finding oxidation states

Find the oxidation state of (a) sulfur in H₂SO₄; (b) manganese in the permanganate ion MnO₄⁻; (c) nitrogen in the ammonium ion NH₄⁺.

(a) $2(+1) + x + 4(-2) = 0 \Rightarrow x = +6$. (b) $x + 4(-2) = -1 \Rightarrow x = +7$. (c) $x + 4(+1) = +1 \Rightarrow x = -3$.

Worked example: is it redox?

For each reaction say whether it is redox, and if so which element is oxidised and which reduced. (a) Fe₂O₃ + 3 CO → 2 Fe + 3 CO₂; (b) NaOH + HCl → NaCl + H₂O.

(a) Fe: +3 → 0, reduced (each atom gains 3 electrons, two atoms gain 6). C: +2 → +4, oxidised (three atoms lose 2 each, so 6 in all). The totals match: 6 and 6. Redox. (b) Na stays +1, O -2, H +1, Cl -1: nothing changes, so it is *not* redox (it is neutralisation).

Common mistakes

- Forgetting that the oxidation states in a compound must sum to zero, then leaving the last one as a guess. Always write the equation $\Sigma = \text{charge}$.
- Using -2 for oxygen in H₂O₂ (it is -1) or +1 for hydrogen in NaH (it is -1). These are the exceptions in the rules.
- Writing 2+ when you mean +2: ions are written with the number first (Mg²⁺), oxidation states with the sign first (+2).
- Concluding that a reaction is redox because it has a gas or a colour change. Only a change in oxidation state counts.

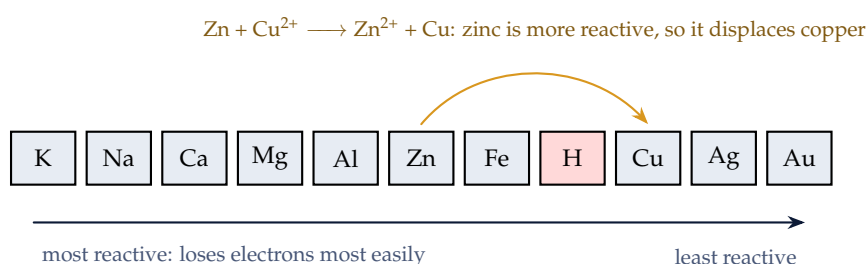
4 Oxidising and reducing agents, displacement and disproportionation (C5.5, C5.6)

Since reduction needs something to take the oxygen (or provide the electrons), we name the species by their *job*:

- An **oxidising agent** takes electrons from another species. It is **itself reduced**.
- A **reducing agent** gives electrons to another species. It is **itself oxidised**.

This is a trap for many students: the oxidising agent is the species that gets *reduced*. In $\text{CuO} + \text{H}_2 \longrightarrow \text{Cu} + \text{H}_2\text{O}$ copper(II) oxide is the oxidising agent (it is reduced and it oxidises the hydrogen), while hydrogen is the reducing agent (it is oxidised). Common oxidising agents are oxygen, chlorine, and acidified potassium manganate(VII); common reducing agents are reactive metals, hydrogen, carbon and carbon monoxide.

The reactivity series and displacement. A more reactive metal is one whose atoms lose electrons more easily; in other words, it is a stronger reducing agent. If a metal is added to a solution of the ions of a less reactive metal, it gives electrons to them and **displaces** that metal:



Zinc displaces copper from copper(II) sulfate, but copper does not displace zinc from zinc sulfate: no reaction takes place if the added metal is the less reactive. The same idea works for halogens: a more reactive halogen displaces a less reactive halide ion, e.g. $\text{Cl}_2 + 2\text{Br}^- \longrightarrow 2\text{Cl}^- + \text{Br}_2$, where chlorine (the oxidising agent) is reduced to chloride and bromide ions (the reducing agent) are oxidised to bromine. Because the reaction is just competition for electrons, reading the *order* of a displacement table is a standard way of finding a reactivity series.

Disproportionation (C5.5). Occasionally one and the same element in a single reaction is both oxidised and reduced. This is **disproportionation**. Chlorine reacts with cold dilute sodium hydroxide:



Chlorine goes from 0 to -1 (in NaCl, reduced) and from 0 to $+1$ (in NaClO, oxidised). The test is simple: look for one element appearing in two products with oxidation states on *either side* of its starting value.

Worked example: agents and disproportionation

- (a) In $2\text{Fe}^{3+} + \text{Sn}^{2+} \longrightarrow 2\text{Fe}^{2+} + \text{Sn}^{4+}$ identify the oxidising and reducing agents. (b) Show that $2\text{H}_2\text{O}_2 \longrightarrow 2\text{H}_2\text{O} + \text{O}_2$ is a disproportionation.
- (a) Fe: $+3 \rightarrow +2$ (reduced), so Fe^{3+} is the oxidising agent; Sn: $+2 \rightarrow +4$ (oxidised), so Sn^{2+} is the reducing agent. (Electrons: $2 \times 1 = 2$ gained, $1 \times 2 = 2$ lost.) (b) Oxygen in H_2O_2 is -1 . In H_2O it is -2 (reduced); in O_2 it is 0 (oxidised). The same element goes both down and up, so it is disproportionation.

Common mistakes

- Mixing up the agents: the oxidising agent is *reduced*, the reducing agent is *oxidised*.
- Using the reactivity series backwards. A metal displaces ions of a metal *below* it in the series, not above it.
- Calling any reaction with two products “disproportionation”. It requires *one element* to go both up and down.

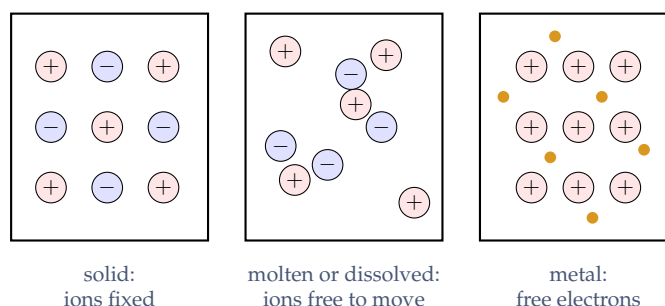
5 Electrolysis: the idea and the apparatus (C12.1 to C12.3)

Some compounds will not decompose just because we want them to: sodium chloride, aluminium oxide and water are very stable. **Electrolysis** uses electrical energy to drive such a reaction. It is the decomposition of an ionic compound, either molten or dissolved in water, by passing an electric current through it. Four words need to be known:

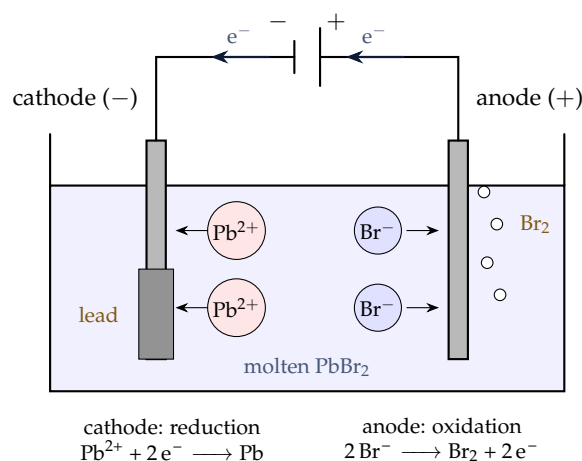
- the **electrolyte**: the molten or dissolved ionic compound, which conducts because its ions are free to move;

- the **electrodes**: the rods or plates (usually graphite or platinum, which are inert) that carry the current into and out of the electrolyte;
- the **cathode**: the **negative** electrode, attached to the negative terminal of the supply;
- the **anode**: the **positive** electrode.

Why it needs an electrolyte with free ions. Solid ionic compounds do not conduct because their ions are locked into the lattice. Once melted or dissolved the ions can move and carry charge. A metal conducts differently: its electrons are free, and no chemical change happens. In electrolysis, the current in the wires is a flow of electrons, but in the electrolyte it is a flow of *ions*.



What happens. The cathode is negative so it attracts positive ions (**cations**), which gain electrons there: *reduction at the cathode*. The anode is positive so it attracts negative ions (**anions**), which lose electrons there: *oxidation at the anode*. The electrons released at the anode travel round the wire, through the power supply, to the cathode, where they are used up. This is why the supply is doing the job of an electron pump. A memory aid for the oxidation and reduction is **RED CAT**, **AN OX**: reduction at the cathode, anode oxidation. Another is **PANIC**: Positive Anode, Negative Is Cathode.



In the picture, molten lead(II) bromide is electrolysed. The lead ions drift to the cathode and each gains two electrons to become a lead atom; the bromide ions drift to the anode, where pairs of them each lose an electron to form a bromine molecule. Because the ions are only ever lost in the ratio of their charges, the number of electrons leaving the anode equals the number arriving at the cathode, which is why the half-equations should have matching electron numbers when you want to compare amounts of products.

Why direct current (C12.2). The supply must be **direct current** (d.c.), which always pushes the same way. With alternating current the polarity changes many times a second, so each electrode would swap between being a cathode and an anode and the products would not separate: any gas or metal formed would immediately be changed back, or mixed together.

Worked example: molten magnesium chloride

Molten magnesium chloride is electrolysed with inert electrodes. State the product at each electrode, write the half-equations, and find the volume of chlorine formed (at rtp, where one mole of gas occupies 24 dm^3) when 2.4 g of magnesium forms. (A_r : $\text{Mg} = 24$.)

Cathode (-): $\text{Mg}^{2+} + 2\text{e}^{-} \longrightarrow \text{Mg}$ (reduction, magnesium metal). Anode (+): $2\text{Cl}^{-} \longrightarrow \text{Cl}_2 + 2\text{e}^{-}$ (oxidation, chlorine gas). Each mole of Mg takes 2 mol of electrons, and each mole of Cl_2 releases 2 mol, so 1 mol of Mg is made along with 1 mol of Cl_2 . Moles of $\text{Mg} = 2.4/24 = 0.10$, so 0.10 mol of Cl_2 : volume = $0.10 \times 24 = 2.4 \text{ dm}^3$.

Common mistakes

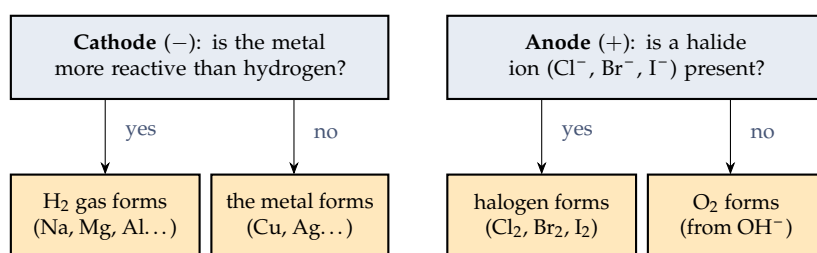
- Putting the products on the wrong electrodes. Metals and hydrogen are positive ions so they go to the *negative* cathode; non-metals (apart from hydrogen) go to the positive anode.
- Saying the electrons flow through the electrolyte. They flow only in the wires; the ions carry the current inside the electrolyte.
- Writing the half-equation for chloride as $\text{Cl}^- \rightarrow \text{Cl} + \text{e}^-$: the gas is a diatomic molecule, so $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$.
- Using alternating current for electrolysis: it must be d.c.

6 Predicting the products (C12.4, C12.5)

Molten ionic compounds. There are only two kinds of ion present, so the answer is simple: the **metal** forms at the cathode and the **non-metal** forms at the anode. Molten sodium chloride gives sodium and chlorine; molten lead(II) bromide gives lead and bromine; molten aluminium oxide gives aluminium and oxygen.

Aqueous solutions. Here there are four kinds of ion, not two, because water itself is slightly ionised: $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$. At each electrode there are two competing ions, and only one is discharged. The deciding rules at GCSE and ESAT level are:

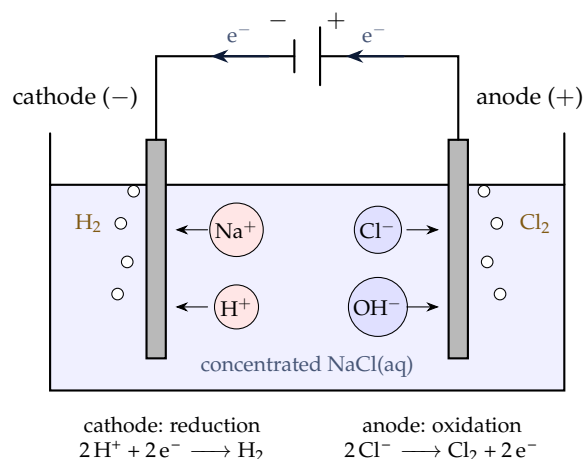
- At the cathode:** if the metal is *more reactive than hydrogen*, hydrogen gas forms ($2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$); if it is *less reactive than hydrogen* (for example copper or silver), the metal is deposited.
- At the anode** (inert electrodes): if a *halide* ion (Cl^- , Br^- , I^-) is present, the halogen forms; if not (for example, sulfate or nitrate solutions), oxygen forms, from hydroxide ions: $4\text{OH}^- \rightarrow 2\text{H}_2\text{O} + \text{O}_2 + 4\text{e}^-$.



molten compound: no water, so the metal forms at the cathode and the non-metal at the anode

Electrolyte (inert electrodes)	Cathode product	Anode product	Left in solution
molten PbBr_2	lead	bromine	—
concentrated NaCl(aq)	hydrogen	chlorine	NaOH(aq)
$\text{CuSO}_4\text{(aq)}$	copper	oxygen	$\text{H}_2\text{SO}_4\text{(aq)}$
dilute $\text{H}_2\text{SO}_4\text{(aq)}$	hydrogen	oxygen	more concentrated acid
$\text{Na}_2\text{SO}_4\text{(aq)}$	hydrogen	oxygen	more concentrated Na_2SO_4

Look at concentrated sodium chloride solution in the next picture. The ions Na^+ and H^+ compete for the cathode; Na^+ is more reactive than hydrogen, so H^+ wins and hydrogen gas bubbles off. At the anode, Cl^- and OH^- compete; a halide is present so chlorine is released. The Na^+ and OH^- ions stay behind: the solution turns into sodium hydroxide, which is why this method is the industrial source of NaOH , Cl_2 and H_2 .



Electrons balance. The same number of electrons passes through each electrode, so products are made in fixed mole ratios. For the electrolysis of dilute sulfuric acid or sodium sulfate, 4 mol of electrons gives $4\text{H}^+ + 4\text{e}^- \longrightarrow 2\text{H}_2$ at the cathode and $4\text{OH}^- \longrightarrow 2\text{H}_2\text{O} + \text{O}_2 + 4\text{e}^-$ at the anode: **two volumes of hydrogen for every one volume of oxygen**. This is just the electrolysis of water, $2\text{H}_2\text{O} \longrightarrow 2\text{H}_2 + \text{O}_2$.

Worked example: predicting aqueous products

Give the products at each inert electrode, and the electrode half-equations, when these are electrolysed: (a) aqueous copper(II) sulfate; (b) aqueous potassium iodide.

- (a) Cathode: copper is less reactive than hydrogen, so $\text{Cu}^{2+} + 2\text{e}^- \longrightarrow \text{Cu}$. Anode: no halide present, so oxygen: $4\text{OH}^- \longrightarrow 2\text{H}_2\text{O} + \text{O}_2 + 4\text{e}^-$. The solution loses its blue colour as copper ions are used up, and becomes acidic.
- (b) Cathode: potassium is more reactive than hydrogen, so hydrogen: $2\text{H}^+ + 2\text{e}^- \longrightarrow \text{H}_2$. Anode: iodide is a halide, so iodine: $2\text{I}^- \longrightarrow \text{I}_2 + 2\text{e}^-$. The solution left is potassium hydroxide.

Exam tip

Write down the ions present (including H^+ and OH^- from water in aqueous solutions), then decide at the cathode: "is the metal above hydrogen?", and at the anode: "is there a halide?". Never guess the product from the formula alone.

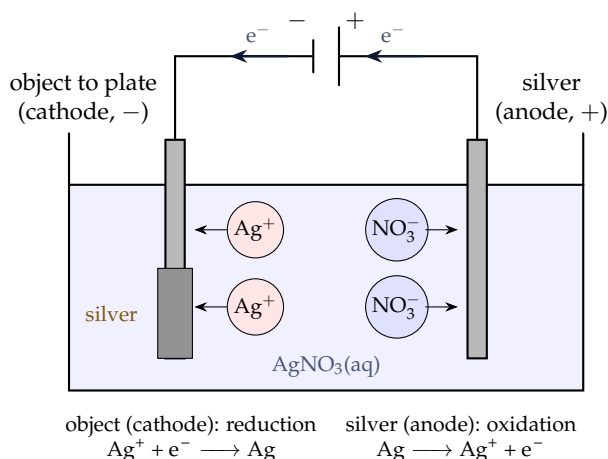
Common mistakes

- Predicting that a reactive metal such as sodium forms at the cathode in an *aqueous* solution. It is hydrogen that forms. Reactive metals are made by electrolysis of the *molten* compound.
- Saying that sulfate or nitrate ions are discharged at the anode. They are not; oxygen forms instead.
- Ignoring concentration: with *dilute* sodium chloride the main anode product is oxygen; the table above is for concentrated solutions. Exam questions say which is meant.
- Forgetting that the electrode may not be inert. With copper electrodes in copper(II) sulfate, the copper anode itself dissolves (next section).

7 Electroplating and extraction of aluminium (C12.6, C12.4)

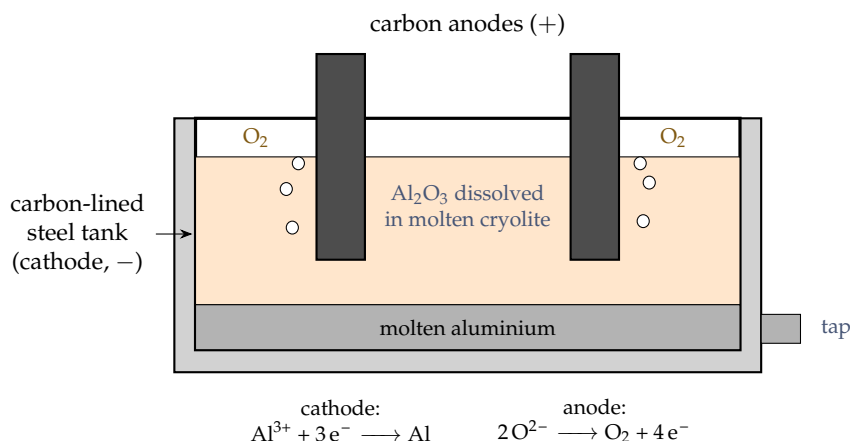
Electroplating coats an object with a thin layer of metal, for decoration (silver on cutlery, gold on jewellery) or protection (chromium on steel to prevent rusting). The set-up is a standard electrolysis cell with three deliberate choices:

- the **object to be plated is the cathode** (so that metal ions are reduced onto it);
- the **anode is a bar of the plating metal**;
- the **electrolyte contains ions of the plating metal**, for example silver nitrate for silver plating.



At the cathode silver ions gain electrons and silver is deposited on the object: $\text{Ag}^+ + \text{e}^- \longrightarrow \text{Ag}$. At the *anode*, silver atoms from the bar lose electrons: $\text{Ag} \longrightarrow \text{Ag}^+ + \text{e}^-$. Because the anode dissolves as fast as the cathode grows, the concentration of silver ions in solution stays the same. The same method, with impure copper as the anode and pure copper as the cathode, is used to purify copper. Since the number of electrons is the same at both electrodes, the mass of metal lost from the anode equals the mass gained at the cathode.

Extraction of aluminium. Aluminium is more reactive than carbon, so its oxide cannot be reduced by heating with carbon. Instead, the oxide is electrolysed. Pure aluminium oxide melts at over 2000°C , which would be very costly, so it is *dissolved in molten cryolite*, a different aluminium compound with a much lower melting point, giving a mixture at about 950°C . Water cannot be used: the product would be hydrogen, not aluminium.



The cell is a steel tank lined with graphite (carbon), which is the cathode. Aluminium ions are reduced there, $\text{Al}^{3+} + 3\text{e}^- \longrightarrow \text{Al}$, to give molten aluminium that sinks to the bottom and is tapped off. Oxide ions are oxidised at the carbon anodes, $2\text{O}^{2-} \longrightarrow \text{O}_2 + 4\text{e}^-$. At that temperature the oxygen reacts with the carbon: $\text{C} + \text{O}_2 \longrightarrow \text{CO}_2$. The anodes gradually burn away and must be replaced regularly.

Worked example: silver plating and aluminium

(a) How many moles of electrons are needed to deposit 2.16 g of silver? (A_r : Ag = 108.) (b) What mass of aluminium can be made from 102 kg of aluminium oxide, Al_2O_3 ? (A_r : O = 16, Al = 27.)

(a) Silver: $2.16/108 = 0.020$ mol. Each Ag^+ needs one electron, so 0.020 mol of electrons. (b) $M_r(\text{Al}_2\text{O}_3) = 2 \times 27 + 3 \times 16 = 102$, so 102 kg is 1000 mol. Each mole contains 2 mol of Al, giving 2000 mol, mass = $2000 \times 27 = 54$ kg.

Common mistakes

- Wiring the object to the positive terminal. The object must be the *cathode* (negative) so that positive metal ions reach it.
- Using a solution with ions of a different metal from the anode; the electrolyte should contain ions of the plating metal.
- Saying cryolite is added to *speed up* the reaction as a catalyst. It dissolves the alumina and lowers the melting

temperature.

- Saying aluminium forms at the anode, or that the cathode burns away. Aluminium forms at the cathode; it is the *carbon anodes* that burn.

Check yourself

Try these without looking back, then compare with the answers on the next page.

1. In $2\text{Mg} + \text{O}_2 \longrightarrow 2\text{MgO}$, which substance is oxidised, in terms of electrons?
2. Find the oxidation state of chromium in $\text{Cr}_2\text{O}_7^{2-}$.
3. Is $\text{AgNO}_3 + \text{NaCl} \longrightarrow \text{AgCl} + \text{NaNO}_3$ a redox reaction?
4. In $\text{Zn} + \text{Cu}^{2+} \longrightarrow \text{Zn}^{2+} + \text{Cu}$, which species is the reducing agent?
5. Name the products at each electrode when molten sodium chloride is electrolysed.
6. What forms at each electrode when dilute sulfuric acid is electrolysed, and in what volume ratio?
7. To silver-plate a spoon, which electrode is the spoon and what is in the solution?
8. Why is aluminium oxide dissolved in molten cryolite before electrolysis?

Answers to the self-check

1. Magnesium: each atom loses two electrons to become Mg^{2+} .
2. $2x + 7(-2) = -2$, so $2x = 12$ and $x = +6$.
3. No: no element changes oxidation state (precipitation).
4. Zinc: it is oxidised ($0 \rightarrow +2$), giving electrons to Cu^{2+} .
5. Sodium at the cathode, chlorine at the anode.
6. Hydrogen at the cathode, oxygen at the anode; 2 : 1 by volume.
7. The cathode; a solution containing silver ions (e.g. silver nitrate), with a silver anode.
8. To lower the temperature needed to melt it, saving energy.

Next step

Heurisco's *ESAT Chemistry: Redox and Electrolysis Practice Set* has twenty original exam-style questions on this lesson, with cell diagrams and a displacement table, in order of difficulty, with a full worked solution and a note on the trap in each.