

Oxidation and reduction (redox): O Level and IGCSE Chemistry notes

These notes cover redox for Cambridge O Level Chemistry 5070 and Cambridge IGCSE Chemistry 0620. They explain the three ways of describing oxidation and reduction, show how to work out oxidation numbers, and apply the ideas to displacement, metal extraction, rusting and electrolysis. Work through the worked examples, then try the Quick check at the end.

1. What redox means

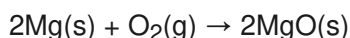
A redox reaction is one in which oxidation and reduction happen at the same time. The word is made from **re**duction and **ox**idation. One substance cannot be oxidised unless another is reduced, because whatever one substance loses (oxygen or electrons) has to be taken by something else.

1.1 Redox in terms of oxygen

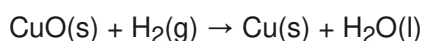
Oxidation is the gain of oxygen by a substance.

Reduction is the loss of oxygen by a substance.

When magnesium burns in air, magnesium gains oxygen, so magnesium is oxidised:



When hydrogen is passed over hot copper(II) oxide, the copper(II) oxide loses oxygen and is reduced to copper. At the same time hydrogen gains oxygen and is oxidised to water:



Worked example 1

In the reaction $\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$, identify the substance oxidised and the substance reduced, using the oxygen definition.

Iron(III) oxide, Fe_2O_3 , loses oxygen to form iron, so iron(III) oxide is **reduced**.

Carbon monoxide, CO , gains oxygen to form carbon dioxide, CO_2 , so carbon monoxide is **oxidised**.

1.2 Redox in terms of electrons

Many redox reactions involve no oxygen at all. For example, sodium burns in chlorine to form sodium chloride. To describe these reactions we look at electrons.

Oxidation is the loss of electrons.

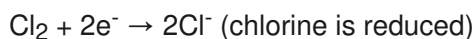
Reduction is the gain of electrons.

Memory aid: **OIL RIG**, Oxidation Is Loss, Reduction Is Gain (of electrons).

In the formation of sodium chloride, each sodium atom loses one electron and each chlorine atom



gains one electron:



1.3 Redox in terms of oxidation numbers

An oxidation number (or oxidation state) shows how many electrons an atom has lost or gained, or has a share in losing or gaining, compared with the free element. Positive means electrons lost; negative means electrons gained.

Oxidation is an increase in oxidation number.

Reduction is a decrease in oxidation number.

The rules for assigning oxidation numbers are:

1. The oxidation number of an atom in an uncombined element is 0. For example, Na, Fe, O₂, Cl₂ and S all have oxidation number 0.
2. The oxidation number of a simple ion equals the charge on the ion. Na⁺ is +1, Mg²⁺ is +2, Cl⁻ is -1, O²⁻ is -2.
3. In a compound the oxidation numbers of all the atoms add up to zero.
4. In a compound ion the oxidation numbers add up to the charge on the ion. For example, in SO₄²⁻ they add up to -2.
5. In most compounds oxygen is -2 and hydrogen is +1. Group I metals are always +1 and Group II metals are always +2 in their compounds. Fluorine is always -1.

Roman numerals in names give an oxidation number directly. Iron(II) sulfate contains iron with oxidation number +2. Iron(III) chloride contains iron with oxidation number +3. Potassium manganate(VII) contains manganese with oxidation number +7.

Worked example 2: finding an oxidation number

Find the oxidation number of manganese in KMnO₄.

Potassium is +1. Each oxygen is -2, so four oxygens give 4 × (-2) = -8.

The compound is neutral, so: (+1) + Mn + (-8) = 0

Mn = +7. This is why the name is potassium manganate(VII).

Worked example 3: oxidation number in an ion

Find the oxidation number of chromium in the dichromate(VI) ion, Cr₂O₇²⁻.

Seven oxygens give 7 × (-2) = -14. Let each chromium be x.

2x + (-14) = -2, so 2x = +12 and x = +6.

Worked example 4: using oxidation numbers to identify redox

Use oxidation numbers to show that the following is a redox reaction: 2FeCl₂ + Cl₂ → 2FeCl₃

Iron: +2 in FeCl_2 and +3 in FeCl_3 . The oxidation number increases, so iron(II) is **oxidised**.

Chlorine in Cl_2 : 0. Chlorine in FeCl_3 : -1. The oxidation number decreases, so chlorine is **reduced**.

Because one oxidation number goes up and another goes down, this is a redox reaction.

Not every reaction is a redox reaction. In neutralisation, such as $\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$, every element keeps the same oxidation number (Na +1, O -2, H +1, Cl -1), so there is no oxidation or reduction. The same is true of precipitation reactions and the reaction of a carbonate with an acid.

2. Oxidising agents and reducing agents

A common source of lost marks is confusing the substance that is oxidised with the oxidising agent. They are opposites.

An **oxidising agent** is a substance that oxidises another substance and is itself reduced during the reaction.

A **reducing agent** is a substance that reduces another substance and is itself oxidised during the reaction.

In terms of electrons, an oxidising agent accepts electrons and a reducing agent gives electrons away. In terms of oxidation numbers, the oxidation number of the oxidising agent goes down and that of the reducing agent goes up.

	Oxidising agent	Reducing agent
Electrons	Gains electrons	Loses electrons
Oxidation number	Decreases	Increases
What happens to it	Is reduced	Is oxidised
Common examples	Oxygen, chlorine, acidified potassium manganate(VII), acidified potassium dichromate(VI), hydrogen peroxide	Hydrogen, carbon, carbon monoxide, reactive metals, potassium iodide, sulfur dioxide

Worked example 5

Zinc reacts with copper(II) ions: $\text{Zn(s)} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu(s)}$. Identify the oxidising agent and the reducing agent.

Zinc goes from 0 to +2: it loses electrons and is oxidised. Zinc is therefore the **reducing agent**.

Copper goes from +2 to 0: it gains electrons and is reduced. Cu^{2+} ions are therefore the **oxidising agent**.

Exam tip: when asked to name the oxidising agent, give the substance as it appears on the left of the equation, for example "copper(II) ions" or " Cu^{2+} ", not "copper".

3. Ionic half-equations

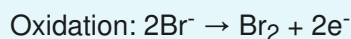
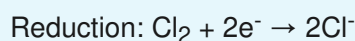
A half-equation shows either the oxidation or the reduction part of a redox reaction on its own, with electrons written in. Adding the two half-equations together, once the electrons balance, gives the overall ionic equation.

Writing half-equations

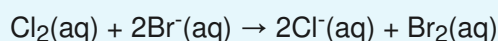
1. Write the species before and after the change.
2. Balance the atoms.
3. Add electrons to the side that needs them so the total charge is the same on both sides.
4. Electrons on the left mean reduction. Electrons on the right mean oxidation.

Worked example 6: combining half-equations

Chlorine displaces bromine from potassium bromide solution. Write the two half-equations and the ionic equation.

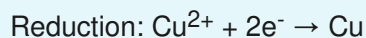
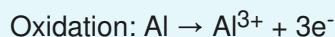


Both involve two electrons, so add them and cancel the electrons:

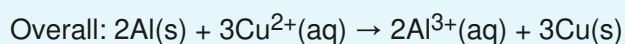
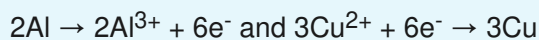


Worked example 7: balancing electrons

Aluminium reacts with copper(II) ions. Write the overall ionic equation.



The lowest common multiple of 3 and 2 is 6. Multiply the first by 2 and the second by 3:



Check: charge on the left is $3 \times (+2) = +6$, charge on the right is $2 \times (+3) = +6$.

4. Tests for oxidising and reducing agents

Two colour-change tests are named in the syllabus. You must know the reagent, the colour change and what the result shows.

4.1 Acidified potassium manganate(VII): test for a reducing agent

Acidified aqueous potassium manganate(VII) is a strong oxidising agent. Its solution is purple because of the MnO_4^- ion. When it oxidises another substance, the manganese is reduced from +7 to +2. Mn^{2+} ions are very pale pink, which looks colourless in dilute solution.

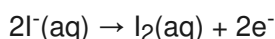
Test for a reducing agent: add acidified aqueous potassium manganate(VII). If a reducing agent

is present, the colour changes from **purple to colourless**.

Iron(II) ions and sulfur dioxide are common examples: Fe^{2+} is oxidised to Fe^{3+} .

4.2 Aqueous potassium iodide: test for an oxidising agent

Iodide ions are easily oxidised to iodine. A colourless solution of potassium iodide therefore acts as a reducing agent and can be used to detect oxidising agents.



Test for an oxidising agent: add aqueous potassium iodide. If an oxidising agent is present, the colour changes from **colourless to brown** (iodine is formed).

Chlorine water, bromine water, hydrogen peroxide, acidified potassium manganate(VII) and iron(III) ions all oxidise iodide ions to brown iodine.

Exam tip: do not mix the two tests up. Manganate(VII) is an oxidising agent, so it *detects reducing agents*. Iodide is a reducing agent, so it *detects oxidising agents*. Always give the colour before and after.

5. Redox in displacement reactions

5.1 Metal displacement

A more reactive metal displaces a less reactive metal from a solution of its salt, or from its oxide when heated. The more reactive metal loses electrons more easily, so it is oxidised and acts as the reducing agent. The ions of the less reactive metal gain electrons and are reduced.

Magnesium in copper(II) sulfate solution: $\text{Mg}(\text{s}) + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{Cu}(\text{s})$. A pink-brown solid forms and the blue colour of the solution fades.

The thermite reaction is a displacement from an oxide: $2\text{Al}(\text{s}) + \text{Fe}_2\text{O}_3(\text{s}) \rightarrow \text{Al}_2\text{O}_3(\text{s}) + 2\text{Fe}(\text{l})$. Aluminium gains oxygen and is oxidised. Iron(III) oxide loses oxygen and is reduced.

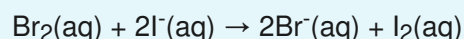
5.2 Halogen displacement

Reactivity in Group VII decreases down the group. A more reactive halogen displaces a less reactive one from a solution of its halide. The more reactive halogen is the stronger oxidising agent: it is reduced to halide ions, and the halide ions of the less reactive halogen are oxidised to the halogen.

Halogen added	Potassium chloride	Potassium bromide	Potassium iodide
Chlorine	No reaction	Orange-brown bromine forms	Brown iodine forms
Bromine	No reaction	No reaction	Brown iodine forms
Iodine	No reaction	No reaction	No reaction

Worked example 8

Bromine water is added to potassium iodide solution. Write the ionic equation and explain, in terms of electrons, why this is a redox reaction.



Each bromine atom gains one electron to become a bromide ion, so bromine is reduced. Each iodide ion loses one electron, and pairs of iodine atoms join to form I_2 , so iodide is oxidised.

Reduction and oxidation take place together, so this is redox. Bromine is the oxidising agent and iodide ions are the reducing agent.

6. Redox in the extraction of metals

Metals are found in ores as compounds, usually oxides, sulfides or carbonates, in which the metal has a positive oxidation number. Extracting the metal always means **reducing** the metal ions to atoms (oxidation number 0). The method depends on the reactivity of the metal.

- Metals above carbon in the reactivity series (for example aluminium, magnesium, sodium) are extracted by electrolysis of a molten compound.
- Metals below carbon (for example zinc and iron) can be extracted by heating the oxide with carbon or carbon monoxide, which act as reducing agents.

6.1 Iron in the blast furnace

Hematite (mainly Fe_2O_3), coke (carbon) and limestone are added at the top. Hot air is blown in near the bottom.

1. Coke burns in the hot air: $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$. Carbon is oxidised. This reaction is strongly exothermic and heats the furnace.
2. Carbon dioxide reacts with more coke: $\text{CO}_2 + \text{C} \rightarrow 2\text{CO}$. Carbon dioxide is reduced and carbon is oxidised.
3. Carbon monoxide reduces the iron(III) oxide: $\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$. Iron goes from +3 to 0.

6.2 Aluminium

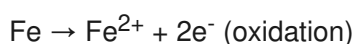
Aluminium is extracted by electrolysis of aluminium oxide dissolved in molten cryolite. At the cathode, aluminium ions are reduced: $\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$. At the anode, oxide ions are oxidised: $2\text{O}^{2-} \rightarrow \text{O}_2 + 4\text{e}^-$. The oxygen reacts with the hot carbon anodes, so the anodes burn away and must be replaced.

7. Redox in rusting

Rusting is the corrosion of iron and steel. It needs both **water and oxygen**. Rust is hydrated iron(III) oxide, often written $\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$.

A simple word equation is: iron + water + oxygen $\rightarrow$ hydrated iron(III) oxide.

Rusting is a redox reaction. Iron atoms lose electrons and are oxidised, first to Fe^{2+} and then to Fe^{3+} . Oxygen gains electrons and is reduced. The iron is the reducing agent and oxygen is the oxidising agent.



7.1 Preventing rusting

- **Barrier methods** keep out water and oxygen: painting, greasing or oiling, coating with plastic, and electroplating with another metal such as tin or chromium.
- **Galvanising** coats iron with zinc. Zinc acts as a barrier, and if the coating is scratched it still protects the iron, because zinc is more reactive.
- **Sacrificial protection** attaches blocks of a more reactive metal, such as zinc or magnesium, to the iron (for example on ship hulls and pipelines).

Sacrificial protection is a redox explanation. The more reactive metal loses electrons more readily than iron, so it is oxidised instead of the iron: $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$. Electrons flow into the iron, so any iron ions that form are reduced back, and the iron is not oxidised. The zinc corrodes away and has to be replaced from time to time.

Exam tip: when explaining sacrificial protection, say that the more reactive metal "loses electrons more easily" or "is oxidised in preference to iron". Simply saying "zinc is more reactive" is usually not enough for full marks.

8. Redox in electrolysis

Electrolysis is the breakdown of an ionic compound, molten or in aqueous solution, by passing an electric current through it. Every electrolysis is a redox process.

At the **cathode** (negative electrode), positive ions gain electrons: **reduction** happens at the cathode.

At the **anode** (positive electrode), negative ions lose electrons: **oxidation** happens at the anode.

Memory aid: **An Ox, Red Cat**, anode oxidation, reduction cathode.

Electrolyte	Cathode half-equation (reduction)	Anode half-equation (oxidation)
Molten lead(II) bromide	$\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$	$2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$
Concentrated aqueous sodium chloride	$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$	$2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$
Dilute sulfuric acid	$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$	$4\text{OH}^- \rightarrow \text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^-$
Aqueous copper(II) sulfate with copper electrodes	$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	$\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$

Worked example 9

During the electrolysis of molten lead(II) bromide, explain which ion is oxidised and which is reduced.

Lead(II) ions move to the cathode and each gains two electrons to form lead atoms. Gain of electrons is reduction, so Pb^{2+} is reduced (oxidation number +2 to 0).

Bromide ions move to the anode and each loses one electron. Pairs of bromine atoms form Br_2 . Loss of electrons is oxidation, so Br^- is oxidised (oxidation number -1 to 0).

9. Summary

	Oxidation	Reduction
Oxygen	Gain of oxygen	Loss of oxygen
Electrons	Loss of electrons	Gain of electrons
Oxidation number	Increase	Decrease
In electrolysis	At the anode	At the cathode
Agent	Caused by an oxidising agent	Caused by a reducing agent

Quick check

1. State the meaning of the term *oxidation* in terms of (a) oxygen and (b) electrons. [2]
2. Copper(II) oxide is heated with carbon: $2\text{CuO} + \text{C} \rightarrow 2\text{Cu} + \text{CO}_2$. Identify the substance that is reduced and the reducing agent. Explain your answer in terms of oxygen. [3]
3. Calculate the oxidation number of (a) sulfur in H_2SO_4 , (b) nitrogen in NO_3^- , (c) chromium in $\text{K}_2\text{Cr}_2\text{O}_7$. [3]
4. State whether each reaction is a redox reaction. Explain using oxidation numbers. (a) $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$ (b) $2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$ [4]
5. Define the term *oxidising agent*. [1]
6. Iron reacts with copper(II) sulfate solution. (a) Write the ionic equation. (b) Write the two half-equations and label each as oxidation or reduction. (c) Name the oxidising agent. [4]
7. Write the ionic equation for the reaction between magnesium and silver ions, Ag^+ . Show how the electrons balance. [3]
8. Describe a test to show that a solution contains a reducing agent. Give the reagent and the colour change. [2]
9. A colourless solution of potassium iodide turns brown when solution X is added. (a) Name the brown substance. (b) What does this show about X? (c) Write the half-equation for the change to the iodide ions. [3]
10. Chlorine water is added to potassium bromide solution. (a) State what is seen. (b) Write the ionic equation. (c) Explain, in terms of electrons, which substance is oxidised. [4]
11. Explain why iodine does not react with potassium chloride solution. [2]
12. In the blast furnace, $\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$. Use oxidation numbers to show that iron(III) oxide is reduced and carbon monoxide is oxidised. [3]
13. Explain why aluminium is extracted by electrolysis and not by heating its oxide with carbon. Write the half-equation for the reaction at the cathode and state whether it is oxidation or reduction. [3]
14. Blocks of zinc are fixed to the steel hull of a ship. Explain, in terms of electrons, how this prevents the hull from rusting. [3]
15. During the electrolysis of concentrated aqueous sodium chloride, chlorine is produced at the anode. Write the half-equation and explain why this is described as oxidation. [2]

Answers

1. (a) Oxidation is the gain of oxygen by a substance. (b) Oxidation is the loss of electrons.
2. Copper(II) oxide loses oxygen to form copper, so copper(II) oxide is reduced. Carbon gains oxygen to form carbon dioxide, so carbon is oxidised. A reducing agent is itself oxidised, so carbon is the reducing agent.
3. (a) H_2SO_4 : $2(+1) + \text{S} + 4(-2) = 0$, so $2 + \text{S} - 8 = 0$ and $\text{S} = +6$.
 (b) NO_3^- : $\text{N} + 3(-2) = -1$, so $\text{N} - 6 = -1$ and $\text{N} = +5$.
 (c) $\text{K}_2\text{Cr}_2\text{O}_7$: $2(+1) + 2\text{Cr} + 7(-2) = 0$, so $2\text{Cr} = 14 - 2 = 12$ and $\text{Cr} = +6$.
4. (a) Not redox. Ca stays +2, C stays +4 (in CaCO_3 : $+2 + \text{C} + 3(-2) = 0$ gives $\text{C} = +4$; in CO_2 : $\text{C} + 2(-2) = 0$ gives $\text{C} = +4$), and O stays -2. No oxidation number changes.
 (b) Redox. Sodium goes from 0 in Na to +1 in NaCl (increase, oxidised). Chlorine goes from 0 in Cl_2 to -1 in NaCl (decrease, reduced).
5. An oxidising agent is a substance that oxidises another substance and is itself reduced.
6. (a) $\text{Fe(s)} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Fe}^{2+}(\text{aq}) + \text{Cu(s)}$
 (b) $\text{Fe} \rightarrow \text{Fe}^{2+} + 2\text{e}^-$ is oxidation (loss of electrons). $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$ is reduction (gain of electrons).
 (c) Copper(II) ions, Cu^{2+} , because they gain electrons and are reduced.
7. Oxidation: $\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$. Reduction: $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$. To balance electrons, multiply the silver half-equation by 2: $2\text{Ag}^+ + 2\text{e}^- \rightarrow 2\text{Ag}$. Adding and cancelling the two electrons: $\text{Mg(s)} + 2\text{Ag}^+(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{Ag(s)}$. Charge check: left +2, right +2.
8. Add acidified aqueous potassium manganate(VII). If a reducing agent is present, the colour changes from purple to colourless.
9. (a) Iodine. (b) X is an oxidising agent, because it has oxidised the iodide ions. (c) $2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^-$
10. (a) The solution turns from colourless (or very pale green) to orange-brown (bromine is formed). (b) $\text{Cl}_2(\text{aq}) + 2\text{Br}^-(\text{aq}) \rightarrow 2\text{Cl}^-(\text{aq}) + \text{Br}_2(\text{aq})$. (c) Bromide ions are oxidised: each Br^- loses one electron to form a bromine atom, and two atoms join to form Br_2 . Loss of electrons is oxidation. (Chlorine gains these electrons and is reduced.)
11. Iodine is less reactive than chlorine, so it is a weaker oxidising agent. It cannot take electrons from chloride ions, so it cannot oxidise chloride ions to chlorine. No displacement takes place.
12. Iron: +3 in Fe_2O_3 ($2\text{Fe} + 3(-2) = 0$, so $\text{Fe} = +3$) and 0 in Fe. The oxidation number decreases, so iron(III) oxide is reduced. Carbon: +2 in CO ($\text{C} + (-2) = 0$) and +4 in CO_2 ($\text{C} + 2(-2) = 0$). The oxidation number increases, so carbon monoxide is oxidised.
13. Aluminium is more reactive than carbon, so carbon cannot remove the oxygen from aluminium oxide (carbon is not a strong enough reducing agent). Electrolysis of aluminium oxide dissolved in molten cryolite is used instead. Cathode: $\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$. This is reduction, because the aluminium ions gain electrons.
14. Zinc is more reactive than iron, so it loses electrons more easily. The zinc is oxidised in preference to the iron: $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$. The electrons released flow into the steel, so the iron does not lose electrons and is not oxidised to iron ions. The zinc corrodes instead (sacrificial protection) and must be replaced when it has worn away.

15. $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$. Each chloride ion loses an electron (oxidation number -1 to 0). Loss of electrons is oxidation.