# Oxidation and Reduction

Matthew Williams • Chemistry • May 15, 2026

## Oxidation and Reduction

Oxidation and reduction always occur together — you cannot have one without the other. This paired process, called a **redox reaction**, underpins rusting, combustion, bleaching, and the reactions that occur at electrodes during electrolysis.

### Oxidation and Reduction in Everyday Life

Redox reactions are not confined to the laboratory. Several familiar processes involve the transfer of electrons:

| Process | Redox involvement |
|---|---|
| Rusting of iron | Iron is oxidised; oxygen is reduced |
| Bleaching of stains | Bleach acts as an oxidising agent, breaking down coloured molecules |
| Browning of cut fruit | Phenolic compounds in fruit are oxidised by oxygen in air |
| Food preservation with SO,or sodium sulfite | These substances act as reducing agents, preventing oxidation of food |

### Definitions of Oxidation and Reduction

There are two complementary sets of definitions, both of which must be known.

#### In Terms of Electrons

| Process | Definition | Memory aid |
|---|---|---|
| Oxidation | Loss of electrons | OIL (Oxidation Is Loss) |
| Reduction | Gain of electrons | RIG (Reduction Is Gain) |

Together: **OIL RIG**

📋 **Example**

In the reaction between zinc and copper(II) sulfate:

$$\mathrm{Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)}$$

Ionic equation:

$$\mathrm{Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)}$$

Zinc: Zn 'Zn²z+ 2e{(loses electrons — **oxidised**)
Copper ions: Cu²z+ 2e{'Cu (gains electrons — **reduced**)

### In Terms of Oxidation Numbers

**Oxidation** is an increase in oxidation number. **Reduction** is a decrease in oxidation number. This definition extends the concept to reactions where discrete electron transfer is less obvious.

## Oxidation Numbers

An **oxidation number** (oxidation state) is the apparent charge an atom would have if all bonds in its compound were fully ionic. It is a bookkeeping tool for tracking electron shifts.

### Rules for Assigning Oxidation Numbers

| Rule | Value |
|---|---|
| Uncombined element (e.g. Na, O, Cl) | 0 |
| Simple monatomic ion | Equals the ionic charge (e.g. Naz= +1, Cl{= −1) |
| Group I metals in compounds | Always +1 |
| Group II metals in compounds | Always +2 |
| Oxygen in compounds | Usually −2 (except in peroxides, where it is −1) |
| Hydrogen in compounds | Usually +1 (except in metal hydrides, where it is −1) |
| Sum of all oxidation numbers in a neutral compound | 0 |
| Sum of all oxidation numbers in a polyatomic ion | Equals the overall charge |

📋 **Example**

Find the oxidation number of manganese in KMnO„

K is +1. O is −2 (×4 = −8). The compound is neutral, so:

$$+1 + x + (-8) = 0 \implies x = +7$$

Manganese is **+7** in potassium manganate(VII).

📋 **Example**

Find the oxidation number of sulfur in SO$_{\text{„}}^{2}${

O is −2 (×4 = −8). The ion has charge 2−, so:

$$x + (-8) = -2 \implies x = +6$$

Sulfur is **+6** in the sulfate ion.

## Identifying Redox Reactions Using Oxidation Numbers

A reaction is a redox reaction if at least one element changes oxidation number. The element whose oxidation number increases is oxidised; the element whose oxidation number decreases is reduced.

📋 **Example**

$$2\mathrm{Fe}^{2+}(\mathrm{aq}) + \mathrm{Cl}_2(\mathrm{g}) \rightarrow 2\mathrm{Fe}^{3+}(\mathrm{aq}) + 2\mathrm{Cl}^{-}(\mathrm{aq})$$

Iron: +2 '+3 (increase — **oxidised**)
Chlorine: 0 '−1 (decrease — **reduced**)

This is a redox reaction.

📋 **Example**

$$\mathrm{CuO}(\mathrm{s}) + \mathrm{H}_2(\mathrm{g}) \rightarrow \mathrm{Cu}(\mathrm{s}) + \mathrm{H}_2\mathrm{O}(\mathrm{l})$$

Copper: +2 '0 (decrease — **reduced**)
Hydrogen: 0 '+1 (increase — **oxidised**)

CuO is reduced; H,is oxidised. This is a redox reaction.

## Oxidising and Reducing Agents

An **oxidising agent** causes oxidation in another substance. To do this, it gains electrons — and so becomes reduced itself.

A **reducing agent** causes reduction in another substance. To do this, it loses electrons — and so becomes oxidised itself.

| Agent | What it does to the other substance | What happens to itself |
|---|---|---|
| Oxidising agent | Oxidises it (removes electrons) | Is itself reduced (gains electrons) |
| Reducing agent | Reduces it (supplies electrons) | Is itself oxidised (loses electrons) |

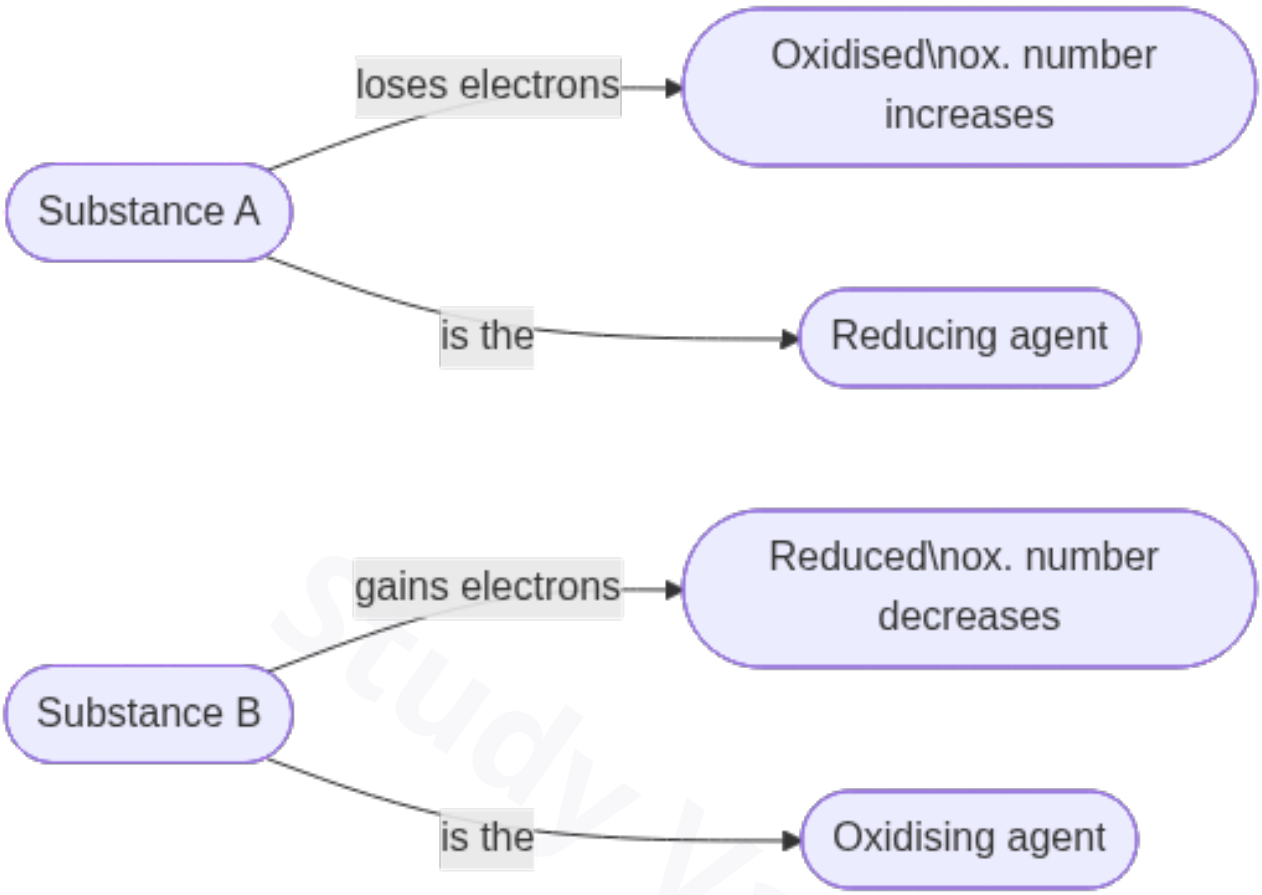
Common oxidising agents: oxygen, chlorine, acidified potassium manganate(VII), hydrogen peroxide, concentrated sulfuric acid.

Common reducing agents: hydrogen, carbon, carbon monoxide, reactive metals (e.g. zinc, iron), iron(II) ions.

> 🧠 **Remember**
> **Hydrogen peroxide** is an example of a substance that can act as both an oxidising agent and a reducing agent, depending on the other reactant. When it reacts with a stronger oxidising agent, it gives up electrons and is oxidised (acts as reducing agent). When it reacts with a stronger reducing agent, it gains electrons and is reduced (acts as oxidising agent).

## Tests for Oxidising and Reducing Agents

| Test | What it detects |
|---|---|
| Add acidified potassium manganate(VII) (purple): decolourises if a reducing agent is present | Reducing agent |
| Add acidified potassium dichromate(VI) (orange): turns green if a reducing agent is present | Reducing agent |
| Add potassium iodide solution: turns brown/black (iodine liberated) if an oxidising agent is present | Oxidising agent |
| Blue litmus paper: turns red in chlorine (oxidising agent decolourises the indicator) | Oxidising agent |




Summary: OIL RIG