

Oxidation Numbers and Redox Reactions

How do you assign oxidation numbers and balance redox equations?

●●○ Intermediate

Redox & Electrochemistry

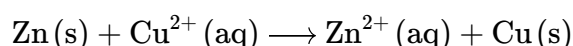
Last reviewed 4 October 2026

Before you start: it helps to understand [Balancing Chemical Equations](#) and [Ionic and Covalent Bonds](#).

What is it?

A **redox reaction** is one in which **electrons are transferred** from one substance to another. It always has two halves:

- **Oxidation:** a loss of electrons. The oxidation number **increases**.
- **Reduction:** a gain of electrons. The oxidation number **decreases**.



Zinc loses 2 electrons (oxidized); each Cu^{2+} ion gains 2 electrons (reduced). Oxidation and reduction always happen **together**: electrons lost by one substance are gained by another.

KEY IDEA

OIL RIG: Oxidation Is Loss, Reduction Is Gain (of electrons). The substance that is oxidized is the **reducing agent**; the substance that is reduced is the **oxidizing agent**.

Why does it matter?

- **Energy.** Batteries, fuel cells, combustion and respiration are all redox reactions.
- **Materials.** Extracting metals from ores, rusting and corrosion protection depend on redox chemistry.
- **Analysis.** Redox titrations (for example with potassium permanganate) measure iron, vitamin C and many other substances.

How does it work?

1. Rules for oxidation numbers

An **oxidation number** is the charge an atom would have if all its bonds were ionic. Apply these rules in order:

1. An **uncombined element** has oxidation number **0** (Zn , O_2 , Cl_2).

2. A **monatomic ion** has an oxidation number equal to its charge (Na^+ : +1; Cl^- : -1).
3. **Fluorine** is always -1. **Group 1** metals are +1 and **group 2** metals are +2 in compounds.
4. **Hydrogen** is usually +1 (but -1 in metal hydrides such as NaH).
5. **Oxygen** is usually -2 (but -1 in peroxides such as H_2O_2).
6. The oxidation numbers add up to **0** in a neutral compound and to the **charge** in an ion.

2. Half-equations

Each half-reaction can be written separately, with electrons shown:

- Oxidation: $\text{Zn} \longrightarrow \text{Zn}^{2+} + 2\text{e}^-$
- Reduction: $\text{Cu}^{2+} + 2\text{e}^- \longrightarrow \text{Cu}$

Electrons appear on the **product** side of an oxidation and the **reactant** side of a reduction.

3. Balancing redox equations in acidic solution

1. Split the reaction into two **half-equations**.
2. Balance atoms other than O and H.
3. Balance **O** by adding H_2O .
4. Balance **H** by adding H^+ .
5. Balance **charge** by adding electrons.
6. Multiply the half-equations so the **electrons lost = electrons gained**, then add them and cancel.
7. Check that **atoms and charge** both balance.

THINK OF IT LIKE THIS

A redox reaction is like a sale: one person can only sell (give) something if another buys (takes) it. The seller is the reducing agent, giving away electrons; the buyer is the oxidizing agent, taking them. The number of items sold must equal the number bought.

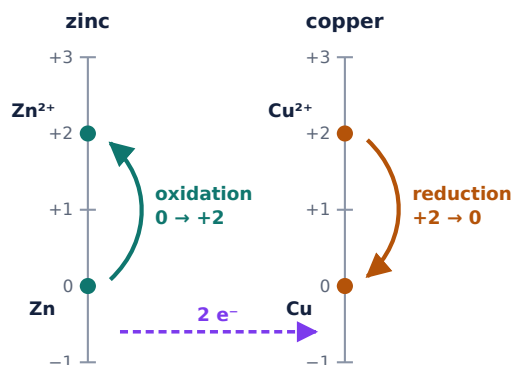
MORE PRECISELY

Oxidation numbers are a bookkeeping tool, not real charges: in covalent molecules the electrons are shared, not transferred. Some reactions are **disproportionations**, in which the same element is both oxidized and reduced; for example, in $2\text{H}_2\text{O}_2 \longrightarrow 2\text{H}_2\text{O} + \text{O}_2$, oxygen goes from -1 to -2 and to 0. In basic solution, balance as in acid and then add OH^- to both sides to neutralize the H^+ .

Visualise it

Oxidation and reduction: $\text{Zn} + \text{Cu}^{2+}$

Oxidation numbers change as electrons move from Zn to Cu^{2+}



Zn is the reducing agent; Cu^{2+} is the oxidizing agent.
Electrons lost = electrons gained (2e^-).

Oxidation raises the oxidation number; reduction lowers it.

Worked example

WORKED EXAMPLE: Assigning oxidation numbers

Question: Find the oxidation number of (a) S in H_2SO_4 (b) Mn in MnO_4^- .

Step 1 (a) H is +1 and O is -2; the compound is neutral: $2(+1) + x + 4(-2) = 0$, so $x = +6$.

Step 2 (b) O is -2; the ion has charge -1: $x + 4(-2) = -1$, so $x = +7$.

WORKED EXAMPLE: Identifying the agents

Question: In $2\text{Al} + 3\text{Cu}^{2+} \rightarrow 2\text{Al}^{3+} + 3\text{Cu}$, which species is oxidized and which is the oxidizing agent?

Step 1 Al: $0 \rightarrow +3$. The oxidation number increases, so **Al is oxidized** (it is the reducing agent).

Step 2 Cu: $+2 \rightarrow 0$. It decreases, so Cu^{2+} is reduced: **Cu^{2+} is the oxidizing agent.**

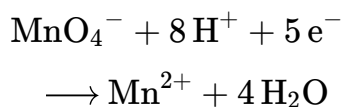
Step 3 Check: 2 Al lose $2 \times 3 = 6$ electrons; 3 Cu^{2+} gain $3 \times 2 = 6$ electrons. ✓

 WORKED EXAMPLE: Balancing in acidic solution

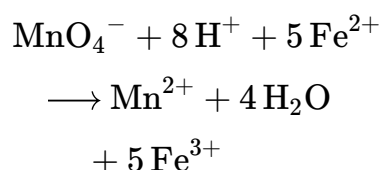
Question: Balance $\text{MnO}_4^- + \text{Fe}^{2+} \longrightarrow \text{Mn}^{2+} + \text{Fe}^{3+}$ in acid.

Step 1 Oxidation: $\text{Fe}^{2+} \longrightarrow \text{Fe}^{3+} + \text{e}^-$

Step 2 Reduction: $\text{MnO}_4^- \longrightarrow \text{Mn}^{2+}$. Add 4 H_2O for oxygen, then 8 H^+ for hydrogen, then 5 electrons for charge:



Step 3 Multiply the oxidation by 5 so that 5 electrons are lost and 5 gained, then add:



Step 4 Check charge: left $-1 + 8 + 10 = +17$; right $+2 + 15 = +17$. ✓

Common mistake

 COMMON MISTAKE: Mixing up oxidation and reduction

Reduction is a **gain** of electrons, even though the oxidation number goes **down**. Use OIL RIG, and remember that the oxidizing agent is itself reduced.

 COMMON MISTAKE: Balancing atoms but not charge

$\text{Fe}^{3+} + \text{Cu} \longrightarrow \text{Fe}^{2+} + \text{Cu}^{2+}$ has balanced atoms but not charge (+3 on the left, +4 on the right). The correct equation is $2\text{Fe}^{3+} + \text{Cu} \longrightarrow 2\text{Fe}^{2+} + \text{Cu}^{2+}$.

 COMMON MISTAKE: Assuming oxygen is always -2

In peroxides such as H_2O_2 , oxygen is -1; in OF_2 it is +2, because fluorine always takes -1.

NOTATION NOTE

- Oxidation numbers are written sign first (+2); ionic charges are written number first (2+).
- Some books use Roman numerals for oxidation numbers in names: iron(III) chloride is FeCl_3 , with Fe at +3.
- “e⁻” stands for an electron.

Remember this

REMEMBER THIS

- OIL RIG: oxidation is loss of electrons (oxidation number up); reduction is gain (oxidation number down).
- Reducing agent = oxidized; oxidizing agent = reduced.
- Oxidation numbers: element 0; ion = charge; F -1; H +1; O -2; sum = overall charge.
- Balance in acid: atoms, then O with H_2O , H with H^+ , charge with e^- ; equalize electrons; check charge.

References

1. Brown, T. L.; LeMay, H. E., Jr.; Bursten, B. E.; Murphy, C. J.; Woodward, P. M.; Stoltzfus, M. W. Chemistry: The Central Science, 15th ed.; Pearson, 2022.

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