

# Galvanic Cells and Electrode Potentials

How does a galvanic cell work and how do you calculate its voltage?

●●○ Intermediate

Redox & Electrochemistry

Last reviewed 4 October 2026

**Before you start:** it helps to understand Oxidation Numbers and Redox Reactions.

## What is it?

A **galvanic** (or voltaic) **cell** uses a spontaneous redox reaction to produce **electricity**. The oxidation and the reduction happen in **separate half-cells**, so the electrons must travel through an external wire, where they can do useful work.

In the **Daniell cell**, a zinc electrode in  $\text{Zn}^{2+}$  solution is connected to a copper electrode in  $\text{Cu}^{2+}$  solution:

- **Anode** (zinc): oxidation,  $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ . It is the **negative** electrode.
- **Cathode** (copper): reduction,  $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$ . It is the **positive** electrode.
- A **salt bridge** lets ions move between the solutions to keep them electrically neutral.

### 💡 KEY IDEA

Electrons flow through the wire from the **anode** (oxidation) to the **cathode** (reduction). The cell voltage is  $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$ ; a **positive**  $E^\circ_{\text{cell}}$  means the reaction is spontaneous.

## Why does it matter?

- **Batteries.** Every battery, from AA cells to lithium-ion packs, is a galvanic cell (or several in series).
- **Corrosion.** Rusting is an unwanted galvanic process; zinc coatings protect steel because zinc is oxidized first.
- **Industry.** Electrolysis (the reverse process, driven by electricity) produces aluminium, chlorine and copper, and electroplates metals.

## How does it work?

### 1. Standard reduction potentials

Each half-cell has a **standard reduction potential**,  $E^\circ$ , measured against the standard hydrogen electrode (defined as 0.00 V) at 25 °C, 1 M and 1 bar. A **more positive**  $E^\circ$  means a

**stronger oxidizing agent** (more easily reduced).

| Half-reaction (reduction)                                                                         | $E^\circ$ (V) |
|---------------------------------------------------------------------------------------------------|---------------|
| $\text{F}_2 + 2\text{e}^- \longrightarrow 2\text{F}^-$                                            | +2.87         |
| $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \longrightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$ | +1.51         |
| $\text{Cl}_2 + 2\text{e}^- \longrightarrow 2\text{Cl}^-$                                          | +1.36         |
| $\text{Ag}^+ + \text{e}^- \longrightarrow \text{Ag}$                                              | +0.80         |
| $\text{Fe}^{3+} + \text{e}^- \longrightarrow \text{Fe}^{2+}$                                      | +0.77         |
| $\text{Cu}^{2+} + 2\text{e}^- \longrightarrow \text{Cu}$                                          | +0.34         |
| $2\text{H}^+ + 2\text{e}^- \longrightarrow \text{H}_2$                                            | 0.00          |
| $\text{Ni}^{2+} + 2\text{e}^- \longrightarrow \text{Ni}$                                          | -0.25         |
| $\text{Zn}^{2+} + 2\text{e}^- \longrightarrow \text{Zn}$                                          | -0.76         |
| $\text{Al}^{3+} + 3\text{e}^- \longrightarrow \text{Al}$                                          | -1.66         |
| $\text{Mg}^{2+} + 2\text{e}^- \longrightarrow \text{Mg}$                                          | -2.37         |

## 2. Cell voltage

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

Use both values **as reduction potentials** from the table, and do **not** multiply  $E^\circ$  by the coefficients: a potential does not depend on how much substance reacts.

## 3. Cell notation

The anode is written on the left and the cathode on the right; a single line | separates phases and a double line || shows the salt bridge:



## 4. Free energy and spontaneity

The cell voltage is linked to the Gibbs free energy change:

$$\Delta G^\circ = -nFE^\circ_{\text{cell}}$$

where  $n$  is the moles of electrons transferred and  $F = 96\,485\text{ C/mol}$  (the **Faraday constant**, the charge on one mole of electrons). Positive  $E^\circ_{\text{cell}}$  gives negative  $\Delta G^\circ$ : a spontaneous reaction.

## 5. Electrolysis

In **electrolysis**, an external power supply **forces** a non-spontaneous redox reaction. The amount of product follows from the charge passed (**Faraday's law**):

$$Q = It \quad n(e^-) = \frac{Q}{F}$$

where  $I$  is the current in amperes ( $A = C/s$ ) and  $t$  is the time in seconds.

### ∞ THINK OF IT LIKE THIS

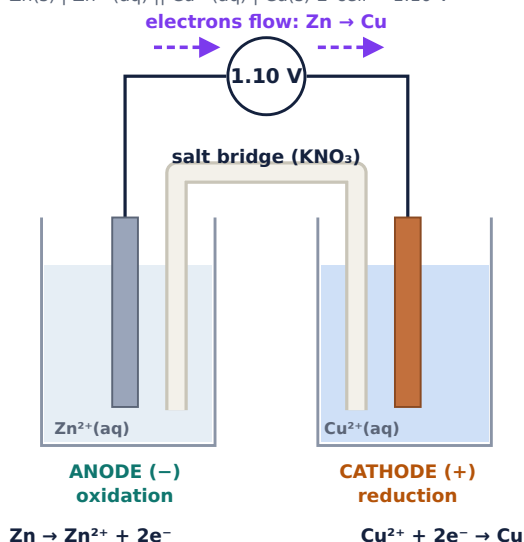
A galvanic cell is like water flowing downhill through a water wheel: the reaction “wants” to happen, and the cell makes it do work on the way. Electrolysis is like pumping water back uphill: it needs an external energy source.

### Q MORE PRECISELY

Standard potentials apply at 1 M concentrations and 25 °C. At other concentrations, the **Nernst equation**,  $E = E^\circ - \frac{RT}{nF} \ln Q$ , gives the actual cell voltage; it falls as the cell runs and  $Q$  increases, reaching zero when the reaction is at equilibrium (a “flat” battery).

## Visualise it

### A galvanic cell: the Daniell cell



Oxidation at the anode, reduction at the cathode, electrons through the wire, ions through the salt bridge.

## Worked example

### WORKED EXAMPLE: Cell voltage and $\Delta G^\circ$

**Question:** A cell is made from  $\text{Ag}^+/\text{Ag}$  and  $\text{Cu}^{2+}/\text{Cu}$  half-cells. Find  $E^\circ_{\text{cell}}$ , write the overall reaction, and calculate  $\Delta G^\circ$ .

**Step 1** The more positive  $E^\circ$  ( $\text{Ag}^+/\text{Ag}$ , +0.80 V) is reduced at the **cathode**; copper (+0.34 V) is oxidized at the **anode**.

**Step 2**  $E^\circ_{\text{cell}} = +0.80 \text{ V} - (+0.34 \text{ V}) = +0.46 \text{ V}$

**Step 3** Overall:  $\text{Cu} + 2 \text{Ag}^+ \longrightarrow \text{Cu}^{2+} + 2 \text{Ag}$ , so  $n = 2$  mol of electrons.

**Step 4** Substitute ( $1 \text{ C} \times 1 \text{ V} = 1 \text{ J}$ ):

$$\begin{aligned}\Delta G^\circ &= -(2 \text{ mol}) \\ &\quad \times (96\,485 \frac{\text{C}}{\text{mol}}) \\ &\quad \times (0.46 \text{ V}) \\ &= -8.9 \times 10^4 \text{ J} \\ &= -89 \text{ kJ}\end{aligned}$$

**Step 5**  $E^\circ_{\text{cell}}$  is positive and  $\Delta G^\circ$  negative: the reaction is **spontaneous**.

### WORKED EXAMPLE: Electroplating copper

**Question:** A current of 2.00 A passes through copper(II) sulfate solution for 30.0 min. What mass of copper is deposited? ( $\text{Cu}^{2+} + 2 \text{e}^- \longrightarrow \text{Cu}$ ;  $\text{Cu} = 63.55 \text{ g/mol}$ )

**Step 1**  $t = 30.0 \text{ min} \times \frac{60 \text{ s}}{1 \text{ min}} = 1800 \text{ s}$

**Step 2**  $Q = It = 2.00 \text{ C/s} \times 1800 \text{ s} = 3600 \text{ C}$

**Step 3**  $n(\text{e}^-) = \frac{3600 \text{ C}}{96\,485 \text{ C/mol}} = 0.03731 \text{ mol}$

**Step 4**  $n(\text{Cu}) = 0.03731 \text{ mol e}^- \times \frac{1 \text{ mol Cu}}{2 \text{ mol e}^-} = 0.01866 \text{ mol}$

**Step 5**  $m = 0.01866 \text{ mol} \times 63.55 \text{ g/mol} = 1.19 \text{ g}$

## Common mistake

### COMMON MISTAKE: Multiplying $E^\circ$ by the coefficients

In  $\text{Cu} + 2 \text{Ag}^+ \longrightarrow \text{Cu}^{2+} + 2 \text{Ag}$ , use  $E^\circ = +0.80 \text{ V}$  for silver, not  $2 \times 0.80 \text{ V}$ . Potentials are intensive: they don't depend on amount.

 **COMMON MISTAKE: Flipping signs twice**

Use  $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$  with both values copied straight from the reduction table. Don't also change the sign of the anode value, or you will subtract twice.

 **COMMON MISTAKE: Mixing up anode and cathode**

In any cell, oxidation happens at the **anode** and reduction at the **cathode** (“**an ox, red cat**”). In a galvanic cell the anode is negative; in electrolysis it is positive.

 **NOTATION NOTE**

- $1 \text{ V} = 1 \text{ J/C}$ , so  $\text{C} \times \text{V} = \text{J}$ ;  $1 \text{ A} = 1 \text{ C/s}$ .
- $F = 96\,485 \text{ C/mol}$  is sometimes rounded to  $96\,500 \text{ C/mol}$ .
- “SHE” = standard hydrogen electrode, the zero of the potential scale.

## Remember this

 **REMEMBER THIS**

- **Galvanic cell:** spontaneous redox makes electricity; anode = oxidation (–), cathode = reduction (+).
- $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$ ; **positive means spontaneous**. Don't multiply  $E^\circ$  by coefficients.
- $\Delta G^\circ = -nFE^\circ_{\text{cell}}$  with  $F = 96\,485 \text{ C/mol}$ .
- **Electrolysis:**  $Q = It$ ,  $n(e^-) = Q/F$ , then use the half-equation's mole ratio.

## 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.

Practise this topic with flashcards and a quiz at [chemistryclarity.com/chemistry/galvanic-cells/](https://chemistryclarity.com/chemistry/galvanic-cells/)