

The Nernst Equation and Cell Potential

How does a cell's voltage change with concentration, and how is it linked to K and ΔG ?

●●○ Intermediate

Redox & Electrochemistry

Last reviewed 6 October 2026

Before you start: it helps to understand [Galvanic Cells and Electrode Potentials](#) and [Chemical Equilibrium and \$K_c\$](#) .

What is it?

Standard cell potentials, E° , apply only when every dissolved species is at **1 mol/L** (and gases at 1 bar). Real cells are rarely standard, and as a battery runs, the concentrations change. The **Nernst equation** gives the actual cell potential E :

$$E = E^\circ - \frac{RT}{nF} \ln Q$$

where n is the number of electrons transferred, $F = 96\,485\text{ C/mol}$ is the Faraday constant and Q is the reaction quotient. At 25°C this becomes:

$$E = E^\circ - \frac{0.0592\text{ V}}{n} \log Q$$

KEY IDEA

As a cell runs, products build up and reactants are used, so **Q increases** and **E falls**. When Q reaches K , the reaction is at equilibrium and $E = 0$: the battery is **flat**.

Why does it matter?

- **Batteries.** The Nernst equation explains why a battery's voltage drops as it is used.
- **Measuring concentrations.** pH meters and ion-selective electrodes measure a voltage that depends on concentration through the Nernst equation.
- **Biology.** Nerve cells hold different ion concentrations inside and outside; the resulting voltage (membrane potential) is a concentration cell.

How does it work?

1. Using the Nernst equation

1. Write the balanced overall equation and find E°_{cell} and n .
2. Write Q (products over reactants; leave out solids).

3. Substitute into $E = E^\circ - \frac{0.0592 \text{ V}}{n} \log Q$.

2. Concentration cells

Two half-cells of the **same** metal in solutions of **different** concentration still give a voltage, even though $E^\circ = 0$. Electrons flow so as to make the concentrations equal: the dilute side is oxidized (anode) and the concentrated side is reduced (cathode).

3. E° , ΔG° and K

Three quantities describe how far a reaction “wants” to go, and each can be found from the others:

$$\Delta G^\circ = -nFE^\circ$$

$$\Delta G^\circ = -RT \ln K$$

$$\log K$$

$$= \frac{nE^\circ}{0.0592 \text{ V}} \quad (\text{at } 25^\circ \text{C})$$

E°_{cell}	ΔG°	K	Reaction at standard conditions
positive	negative	greater than 1	spontaneous, products favoured
zero	zero	1	at equilibrium
negative	positive	less than 1	non-spontaneous, reactants favoured

💡 THINK OF IT LIKE THIS

A cell is like water flowing between two tanks. E° is the height difference when both tanks start at a standard level. As water flows, the upper tank empties and the lower one fills, so the height difference (E) shrinks until the levels are equal and the flow stops (equilibrium, $E = 0$).

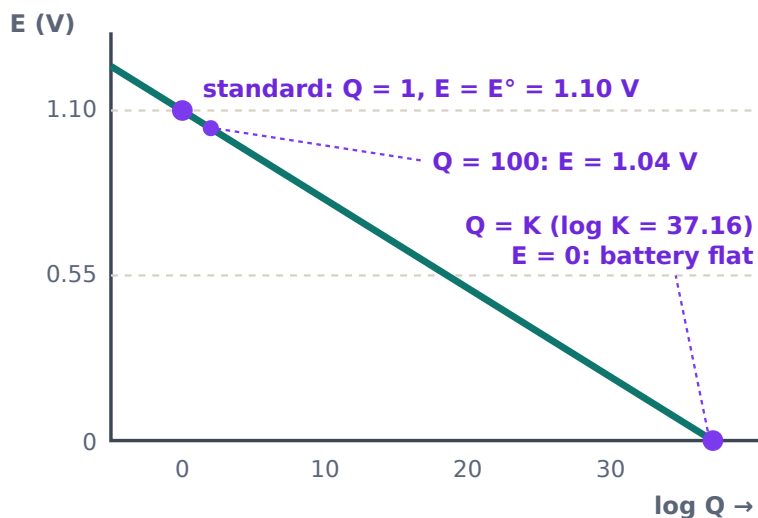
🔍 MORE PRECISELY

The value 0.0592 V comes from $\frac{RT}{F} \ln 10$ at 298.15 K; at other temperatures use the full form. Because $\log K = nE^\circ / 0.0592$, a standard potential of only about 0.2 V for a 2-electron reaction already gives $K \approx 10^7$: small voltages correspond to very large equilibrium constants. Strictly, Q uses activities, so the Nernst equation is most accurate for dilute solutions.

Visualise it

Cell potential falls as Q rises

Zn–Cu cell at 25 °C: $E = 1.10 \text{ V} - (0.0592 \text{ V} / 2) \log Q$



For the Zn–Cu cell, E falls by 0.0296 V for every tenfold increase in Q , reaching 0 when $Q = K$.

Worked example

WORKED EXAMPLE: A non-standard Zn–Cu cell

Question: Calculate E at 25 °C for $\text{Zn(s)} + \text{Cu}^{2+}(\text{aq}) \longrightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu(s)}$ when $[\text{Zn}^{2+}] = 1.0 \text{ mol/L}$ and $[\text{Cu}^{2+}] = 0.010 \text{ mol/L}$.

Step 1 $E^\circ = 0.34 \text{ V} - (-0.76 \text{ V}) = 1.10 \text{ V}$; $n = 2$.

Step 2 $Q = \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} = \frac{1.0}{0.010} = 100$

Step 3 Nernst:

$$\begin{aligned} E &= 1.10 \text{ V} \\ &\quad - \frac{0.0592 \text{ V}}{2} \log 100 \\ &= 1.10 \text{ V} - 0.0592 \text{ V} \\ &= 1.04 \text{ V} \end{aligned}$$

Step 4 Less Cu^{2+} (a reactant) means a slightly lower voltage.

 WORKED EXAMPLE: A concentration cell

Question: Two copper electrodes dip into 0.0010 mol/L and 1.0 mol/L Cu^{2+} . Calculate the cell potential at 25 °C.

Step 1 $E^\circ = 0 \text{ V}$ (same half-reaction); $n = 2$. The dilute side is the anode: $Q = \frac{[\text{Cu}^{2+}]_{\text{dilute}}}{[\text{Cu}^{2+}]_{\text{concentrated}}} = \frac{0.0010}{1.0}$.

Step 2 Nernst:

$$\begin{aligned} E &= 0 - \frac{0.0592 \text{ V}}{2} \log(1.0 \\ &\quad \times 10^{-3}) \\ &= 0.089 \text{ V} \end{aligned}$$

 WORKED EXAMPLE: From E° to K

Question: Calculate K at 25 °C for the Zn–Cu reaction ($E^\circ = 1.10 \text{ V}$, $n = 2$).

$$\begin{aligned} \log K &= \frac{2 \times 1.10 \text{ V}}{0.0592 \text{ V}} \\ &= 37.16 \\ K &= 10^{37.16} = 1.5 \times 10^{37} \end{aligned}$$

The reaction goes essentially to completion.

 WORKED EXAMPLE: ΔG from E

Question: Calculate ΔG for the non-standard Zn–Cu cell above ($E = 1.04 \text{ V}$).

$$\begin{aligned} \Delta G &= -nFE \\ &= -2 \times 96\,485 \frac{\text{C}}{\text{mol}} \times 1.04 \text{ V} \\ &= -2.01 \times 10^5 \text{ J/mol} \\ &= -201 \text{ kJ/mol} \end{aligned}$$

(1 C $\times$ 1 V = 1 J.)

Common mistake

COMMON MISTAKE: Using the wrong n

n is the number of electrons transferred in the **balanced overall** equation: 2 for $\text{Zn} + \text{Cu}^{2+}$, not 1.

COMMON MISTAKE: Getting Q upside down

Q is products over reactants, like K . For $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$, $Q = [\text{Zn}^{2+}]/[\text{Cu}^{2+}]$. Solids (Zn, Cu) are left out.

COMMON MISTAKE: Mixing $\ln$ and $\log$

$\frac{RT}{nF} \ln Q$ and $\frac{0.0592 \text{ V}}{n} \log Q$ are the same at 25 °C. Don't use 0.0592 with $\ln$.

NOTATION NOTE

- E° : standard potential; E : actual potential. $F = 96\,485 \text{ C/mol}$.
- $1 \text{ V} = 1 \text{ J/C}$, so nFE comes out in J/mol.

Remember this

REMEMBER THIS

- Nernst: $E = E^\circ - \frac{RT}{nF} \ln Q$; at 25 °C, $E = E^\circ - \frac{0.0592 \text{ V}}{n} \log Q$.
- As the cell runs, Q rises and E falls; at equilibrium $Q = K$ and $E = 0$.
- Concentration cells: same electrodes, different concentrations, $E^\circ = 0$ but $E \neq 0$.
- $\Delta G^\circ = -nFE^\circ = -RT \ln K$; $\log K = nE^\circ / 0.0592 \text{ V}$ at 25 °C.

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/nernst-equation/