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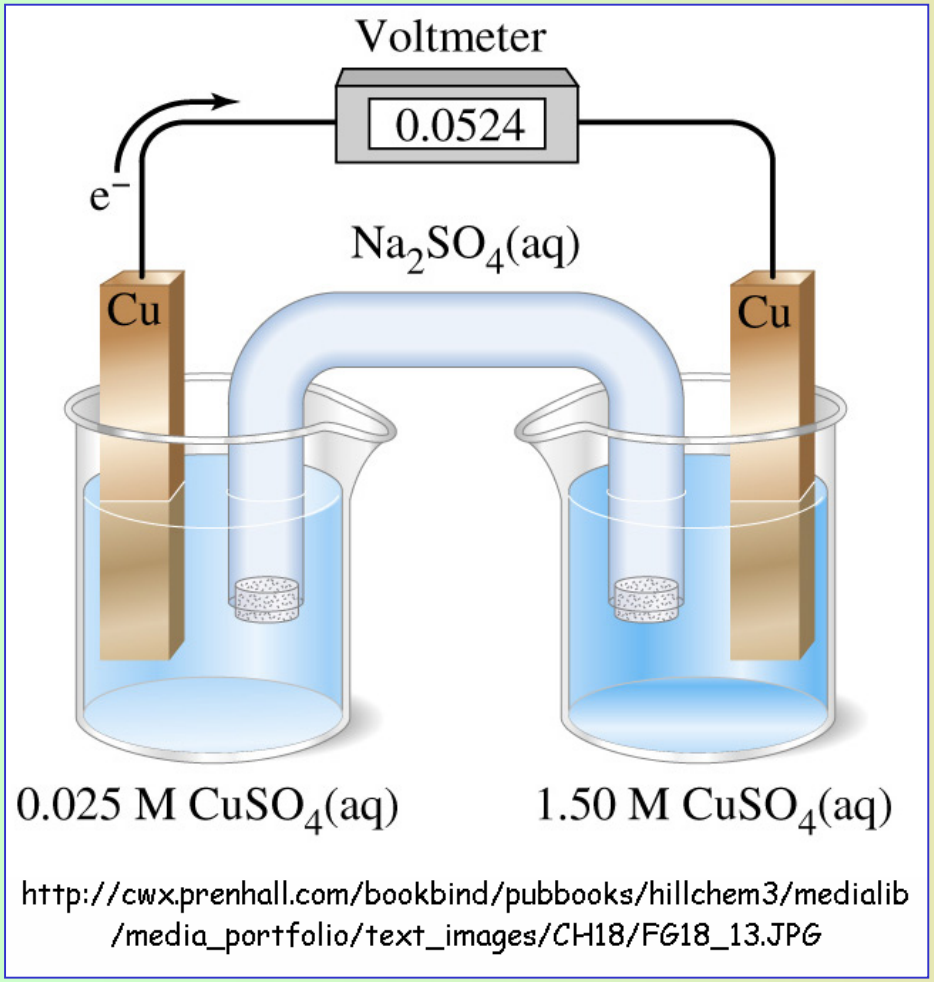
# The Nernst Equation

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## The Nernst Equation

Sometimes the system is under nonstandard conditions. That is, concentrations of dissolved solutes often is not 1 M, and gases are not necessarily at 1 atm.

Cell potentials depend on concentrations and pressure.

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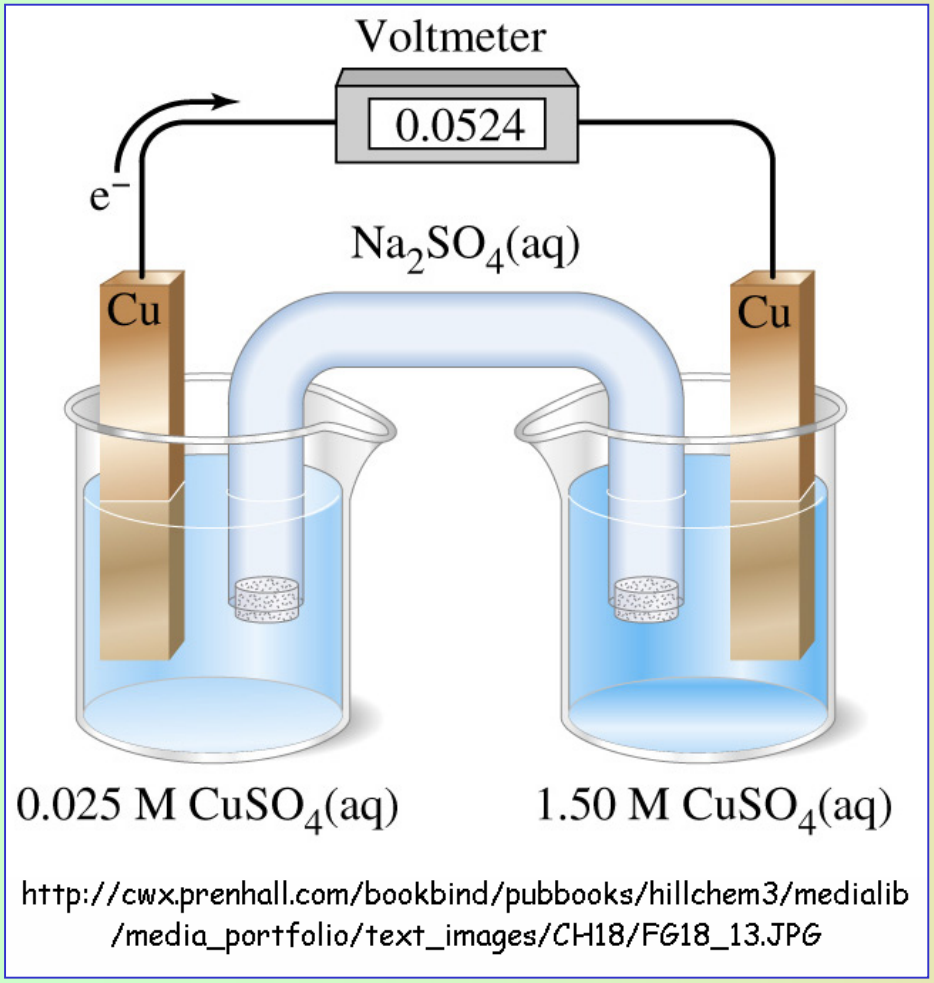


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## The Nernst Equation

The Nernst equation is used to convert between standard cell potentials and potentials of electrochemical cells operating under nonstandard concentration conditions.

$$E = E^0 - \frac{0.059}{n} \log Q$$

E ... cell potential
$E^0$ ... standard cell potential
n ... number of electrons transferred in the reaction
Q ... reaction quotient; contains only the species with concentrations (or pressures)

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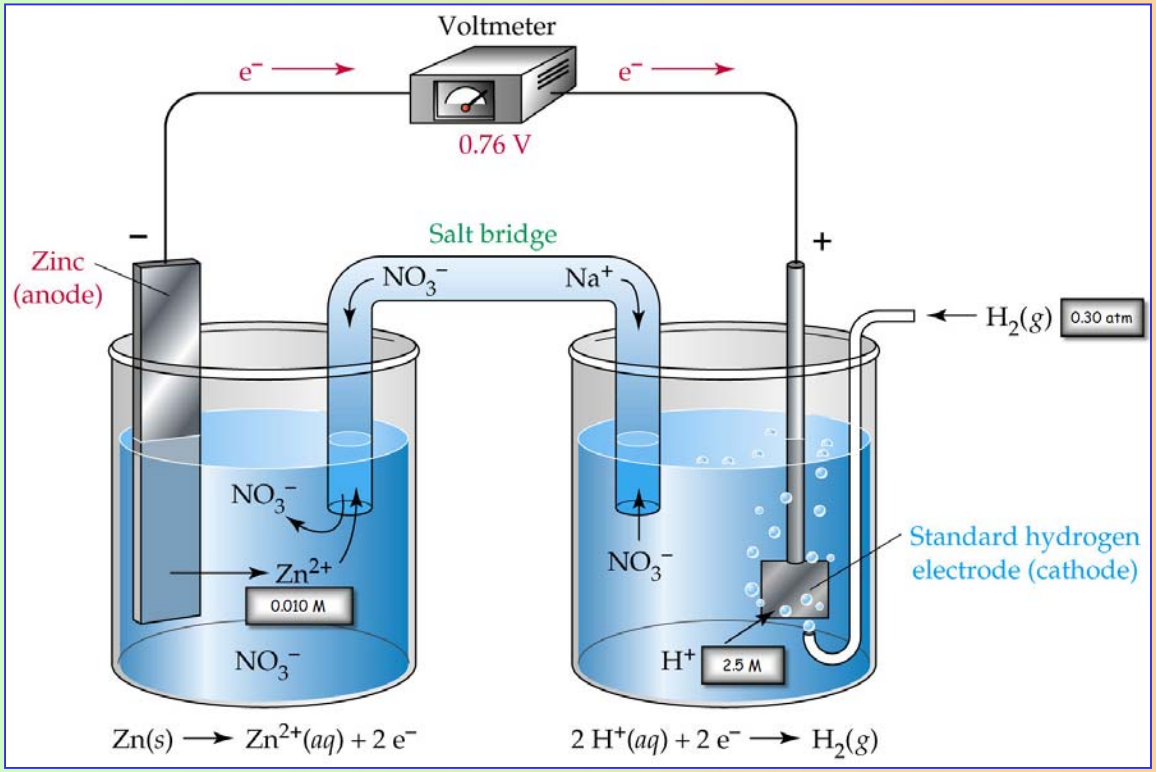
# The Nernst Equation

## Sample Problem

A chemist constructs a voltaic cell consisting of a $Zn/Zn^{2+}$ electrode and the $H_2/H^+$ electrode under the following conditions:

$[Zn^{2+}] = 0.010\ M$ $\quad$ $[H^+] = 2.5\ M$ $\quad$ $P_{H_2} = 0.30\ atm$

Calculate $E_{cell}$ at 25 °C



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# The Nernst Equation

## Solution

Determination of the cell reaction and $E^0_{cell}$:

$$2\ H^+\ (aq) + 2\ e^-\ \rightarrow\ H_2\ (g) \qquad E^0 = 0.00\ V$$

$$Zn\ (s)\ \rightarrow\ Zn^{2+}\ (aq) + 2\ e^- \qquad -E^0 = 0.76\ V$$

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$$2\ H^+\ (aq) + Zn\ (s)\ \rightarrow\ H_2\ (g) + Zn^{2+}\ (aq) \qquad E^0_{cell} = 0.76\ V$$

Determination of $E_{cell}$:

$$E = E^0 - \frac{0.059}{n} \log Q = 0.76 - \frac{0.059}{2} \log \frac{[Zn^{2+}]\ P_{H_2}}{[H^+]^2}$$

$$E = 0.76 - \frac{0.059}{2} \log \frac{0.01 * 0.3}{2.5^2} = 0.86\ V$$