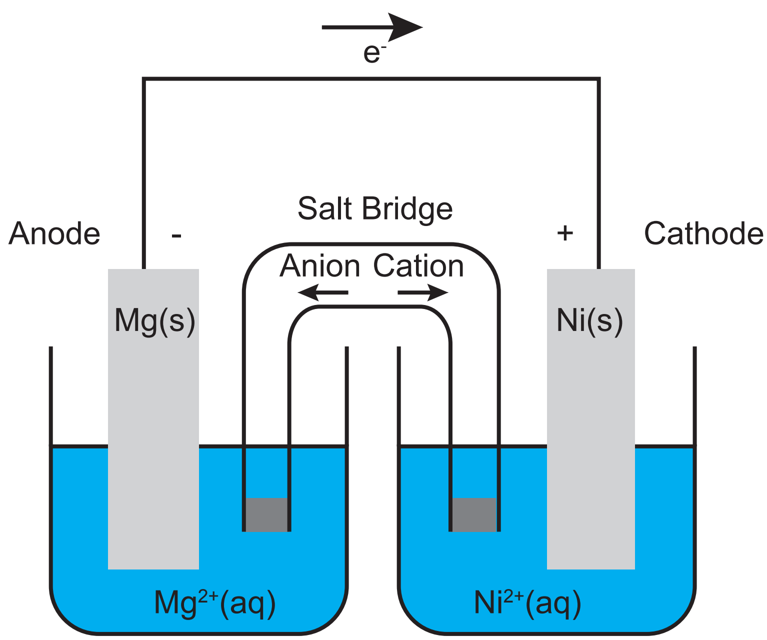
1. Sketch a voltaic cell for the following redox reaction. Label the anode and cathode and indicate the half-reaction that occurs at each electrode and the species present in each solution. Also indicate the direction of electron flow. Calculate the standard cell potential. Write the line notation of the cell.

$$Ni^{2+}(aq) + Mg(s) \longrightarrow Ni(s) + Mg^{2+}(aq)$$

Two half-reactions:
Reduction (cathode): $Ni^{2+}(aq) + 2\,e^- \longrightarrow Ni(s)$
Oxidation (anode): $Mg(s) \longrightarrow Mg^{2+}(aq) + 2\,e^-$



$E^\circ_{cat} = -0.23\ V\ E^\circ_{an} = -2.37\ V$

$E^\circ_{cell} = -0.23 - (-2.37) = 2.14\ V$

$Mg(s) \mid Mg^{2+}(aq, 1\,M) \parallel Ni^{2+}(aq, 1\,M) \mid Ni(s)$

2. Use tabulated electrode potentials to calculate ΔG°_{rxn} and the equilibrium constant, K, for the following reaction at 25 °C.

$$O_2(g) + 2\,H_2O(\ell) + 2\,Cu(s) \longrightarrow 4\,OH^-(aq) + 2\,Cu^{2+}(aq)$$

Oxidation (anode): $2\,Cu(s) \longrightarrow 2\,Cu^{2+}(aq) + 4\,e^-$ $E^\circ_{an} = 0.34\ V$
Reduction (cathode): $O_2(g) + 2\,H_2O(\ell) + 4\,e^- \longrightarrow 4\,OH^-(aq)$ $E^\circ_{cat} = 0.40\ V$

$E^\circ_{cell} = 0.40\ V - 0.34\ V = 0.06\ V$

Remember, $1\ V = 1\ J/C$
Then, $\Delta G^\circ_{rxn} = -nFE^\circ_{cell} = -(4\,mol\ e^-)(96485C/mol\ e^-)(0.06\ J/C) = -23,156\ J = -20\ kJ$

Then, $E^\circ_{cell} = 0.06\ V = \frac{0.0592}{4\ mol\ e^-} \log K \Rightarrow \log K = 4.05 \Rightarrow K = 1 \times 10^4$

3. A voltaic cell employs the redox reaction:

$$2\,Fe^{3+}(aq) + 3\,Mg(s) \longrightarrow 2\,Fe(s) + 3\,Mg^{2+}(aq)$$

Calculate the cell potential at 25 °C under standard conditions and $[Fe^{3+}]$ = 2.00 M and $[Mg^{2+}] = 1.5\times10^{-3}$ M.

Oxidation (anode): $3\,Mg(s) \longrightarrow 3\,Mg^{2+} + 6\,e^-$ $E^\circ_{an} = -2.37\ V$
Reduction (cathode): $2\,Fe^{3+} + 6\,e^- \longrightarrow 2\,Fe(s)$ $E^\circ_{cat} = -0.036\ V$

$E^\circ_{cell} = -0.036\ V - (-2.37\ V) = 2.334\ V$

Then, under non-standard conditions: $E_{cell} = E^\circ_{cell} - \frac{0.0592\ V}{n}\log Q$

In this case, $n = 6$, and $Q = \frac{[Mg^{2+}]^3}{[Fe^{3+}]^2}$

$E_{cell} = 2.334\ V - \frac{0.0592\ V}{6\ mol}\log\left(\frac{(1.5\times10^{-3})^3}{(2.00)^2}\right)$ = 2.42 V

4. A concentration cell consists of two Sn/Sn^{2+} half-cells. The cell has a potential of 0.100 V at 25°C. What is the ratio of the Sn^{2+} concentration in the two half-cells?
The overall reaction is $Sn(s) + Sn^{2+}(aq) \longrightarrow Sn^{2+}(aq) + Sn(s)$

Oxidation (anode): $Sn(s) \longrightarrow Sn^{2+}(aq) + 2\,e^-$
Reduction (cathode): $Sn^{2+}(aq) + 2\,e^- \longrightarrow Sn(s)$

$E_{cell} = E^\circ_{cell} - \frac{0.0592\ V}{n} \log Q$

The $E^\circ_{cell} = 0\ V$, so $0.100V = -\frac{0.0592}{2} \log Q = -(0.0296) \log Q$

$\log Q = -3.38$

$Q = 10^{-3.38} = 0.000417.$

The ratio of $\frac{[Sn^{2+}]_{an}}{[Sn^{2+}]_{cat}} = 4.2 \times 10^{-4}$