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- Calculate the standard free-energy change for the following reaction at 298 K.

$$2Au(s) + 3Mg^{2+}(1.0\ M) \rightarrow 2Au^{3+}(1.0\ M) + 3Mg(s)$$

The half-cell reduction reactions and potentials are:

$$\mathbf{Au^{3+}(aq) + 3e^{-} \rightarrow Au(s) \qquad E^{0} = +1.50\ V}$$
$$\mathbf{Mg^{2+}(aq) + 2e^{-} \rightarrow Mg(s) \qquad E^{0} = -2.36\ V}$$

In the reaction above, the Au is undergoing oxidation so its potential is reversed and the overall cell potential is:

$$\mathbf{E^{0}_{cell} = (-2.36) - (+1.50) = -3.86\ V}$$

Using $\Delta G^{0} = -nFE^{0}$ for this six electron reaction:

$$\mathbf{\Delta G^{0} = -(6) \times (96485) \times (-3.86) = +2.23 \times 10^{6}\ J\ mol^{1} = +2.23 \times 10^{3}\ kJ\ mol^{1}}$$

Answer: **$+2.23 \times 10^{3}$ kJ mol^{-1}**