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

$$2Au(s) + 3Mg^{2+}(1.0 \text{ M}) \rightarrow 2Au^{3+}(1.0 \text{ M}) + 3Mg(s)$$
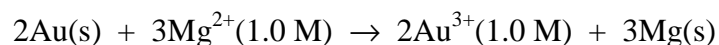
The half-cell reduction reactions and potentials are:

$Au^{3+}(aq) + 3e^- \rightarrow Au(s)$ $E^0 = +1.50 \text{ V}$
$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$ $E^0 = -2.36 \text{ V}$
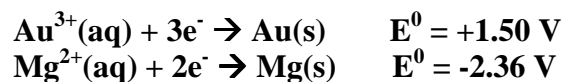
In the reaction above, the Au is undergoing oxidation so its potential is reversed and the overall cell potential is:

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

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

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

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