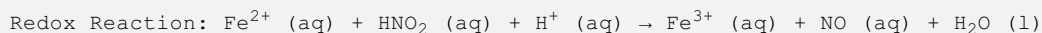


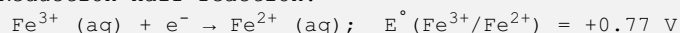
Electrochemistry – Calculating Gibbs Free Energy

Given Data:



Step 1: Determine the half-reactions and their standard reduction potentials (E°) from standard tables.

1. Reduction half-reaction:



2. Oxidation half-reaction:



Explanation: This step involves finding the standard reduction potentials (E°) for the involved half-reactions from a standard table.

Supporting Statement: To calculate Gibbs Free Energy change, the standard reduction potentials for each half-reaction must be identified from standard tables or verified sources.

Step 2: Determine the overall standard cell potential (E°_{cell}).

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

Given:

$$E^\circ_{\text{cathode}} = +0.77 \text{ V}$$

$$E^\circ_{\text{anode}} = +0.994 \text{ V}$$

$$E^\circ_{\text{cell}} = 0.77 \text{ V} - 0.994 \text{ V} = -0.224 \text{ V}$$

Explanation: The standard cell potential (E°_{cell}) is calculated by subtracting the anode potential from the cathode potential. This reflects the overall potential difference driving the redox reaction.

Supporting Statement: Using the given standard reduction potentials, the overall cell potential can be calculated to proceed with Gibbs Free Energy calculations.

Step 3: Calculate the standard free energy change (ΔG°).

Using the formula:

$$\Delta G^\circ = -nFE^\circ_{\text{cell}}$$

Where:

- n = number of moles of electrons transferred (1 mol here, as only one electron is involved in both half-reactions)
- F = Faraday's constant (96485 C/mol)
- E°_{cell} = standard cell potential

Substitute the values:

$$\Delta G^\circ = - (1 \text{ mol}) \times (96485 \text{ C/mol}) \times (-0.224 \text{ V})$$

$$\Delta G^\circ = 21612.64 \text{ J/mol}$$

Explanation: The standard free energy change (ΔG°) is determined using the relationship $\Delta G^\circ = -nFE^\circ_{\text{cell}}$. This indicates the thermodynamic favorability of the reaction.

Supporting Statement: The equation for Gibbs Free Energy relies on the number of moles of electrons transferred, Faraday's constant, and the standard cell potential.

Final Step: Report the final answer with appropriate significant digits.

$$\Delta G^\circ \approx 21612 \text{ J/mol}$$

Explanation: The final value is adjusted to the appropriate significant digits considering the precision of given constants and measured values.

Supporting Statement: Ensure the final result is presented with correct significant figures based on input data precision.

Final Solution:

The standard reaction free energy (ΔG°) for the given redox reaction is approximately 21612 J/mol.