CONCEPT: CELL POTENTIAL AND EQUILIBRIUM

- **Gibbs Free Energy** (____) is the bridge to the standard cell potential (_____) and the equilibrium constant (___).
  - □ This connection can be seen as:

E°_{Cell} , ΔG, K Formulas

E°_{cell} – ΔG Formula

ΔG° = – ___ • ___ • _____

- □ ΔG° = Gibbs Free Energy in _____.
- □ n = moles of e^- transferred
- □ F = Faraday's Constant in $96{,}485 \ \frac{C}{mol\ e^-}$
- □ E°_{cell} = Cell Potential in _____ (V)

E°_{cell} – K Formula

– _____ • _____ • _____ = – _____ • _____ • _____

K – ΔG Formula

ΔG° = – ___ • ___ • _____

- □ R = Gas Constant in $8.314 \ \frac{J}{mol \bullet K}$
- □ T = Temperature in ___
- □ K = Equilibrium Constant

  - □ Isolating the standard cell potential gives the formula:

E°_{Cell} – K Formula (Simplified)

E°_{Cell} = _____ ln ___

EXAMPLE: A certain electrochemical reaction involves the transferring of 4 electrons. If the value of its equilibrium constant is 2.50×10^{12} at 25.0ºC, calculate the standard cell potential.

PRACTICE: Calculate the equilibrium constant for the following reaction at 25ºC.

$$Fe\ (s) + I_2\ (s) \longrightarrow Fe^{2+}\ (aq) + 2\ I^-\ (aq)$$

Given the following reduction potentials:

$Fe^{2+}\ (aq) + 2\ e^- \longrightarrow Fe\ (s)$ $\qquad E^{\circ}_{red} = -0.45\ V$

$I_2\ (s) + 2\ e^- \longrightarrow 2\ I^-\ (aq)$ $\qquad E^{\circ}_{red} = +0.54\ V$