

April 15, 2002

➤ Exam #3

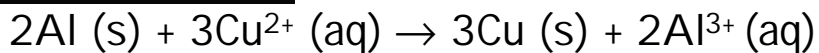
✓1999 Exam #3 Questions - answers posted

✓Exam Info Page - updated!

✓**Monday, 4/15 Review/Problem Session: 3 - 4 pm, A531 Cook**

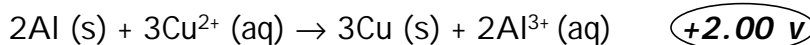
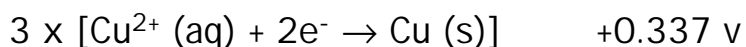
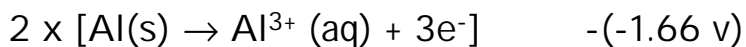
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Example



E°_{cell} ? ΔG° ? K ?

E°_{cell} :



reaction is *spontaneous*

2

Now for ΔG°

$$\Delta G^\circ = -n F E^\circ_{\text{cell}}$$

$$\Delta G^\circ = -(6 \text{ mol e}^-/\text{mol})(9.6487 \times 10^4 \text{ C/mol e}^-)(2.00 \text{ v})$$

$$\Delta G^\circ = -1.158 \times 10^6 \text{ C}\cdot\text{v}/\text{mol}$$

Joules

$$\Delta G^\circ = \underline{-1.16 \times 10^3 \text{ kJ/mol}}$$

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Equilibrium at last

➤ Two paths:

- From $\Delta G^\circ (= -RT \ln K)$
- From $E^\circ_{\text{cell}} = (0.0592/n) \text{Log} K$

$$\text{Log} K = \frac{(E^\circ_{\text{cell}})n}{0.0592}$$

$$\text{Log} K = \frac{(2.00 \text{ v})(6 \text{ mol e}^-)}{0.0592} = 202.703$$

$$K = 10^{202.703} = 5.043 \times 10^{203} = \underline{10^{203}}$$

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Non-Standard States?

Recall:

$$\mathbf{DG} = \mathbf{DG}^{\circ} + \mathbf{RTLnQ}$$

Since:

$$\Delta G = -n F E_{\text{cell}}$$

Substituting:

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{RT}{nF} \text{LnQ}$$

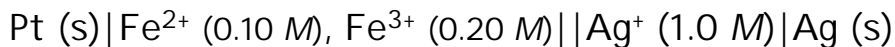
The Nernst Equation

At 298.15 K:

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - (0.0592/n)\text{LogQ}$$

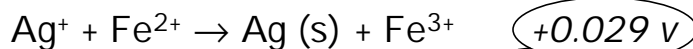
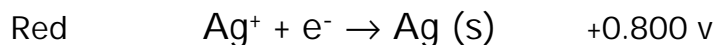
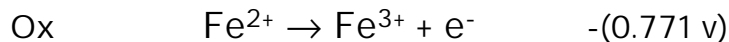
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Nernst Example



Calculate E_{cell}

First, determine E°_{cell} :



E°_{cell}

6

Example Continues

Next, find value of Q :

$$Q = \frac{[\text{Fe}^{3+}]}{[\text{Ag}^+][\text{Fe}^{2+}]} = \frac{(0.20)}{(1.0)(0.10)} = 2.0$$

Lastly, into the *Nernst Equation*:

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - (0.0592/n)\text{Log}Q$$

$$E_{\text{cell}} = 0.029 - (0.0592/1)\text{Log}(2.0)$$

$$E_{\text{cell}} = 0.029 - 0.01782 = 0.01179 \text{ volts}$$

$$E_{\text{cell}} = \underline{\underline{0.012 \text{ volts}}}$$

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Nernst (huh!)

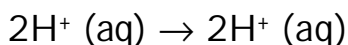
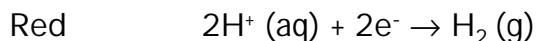
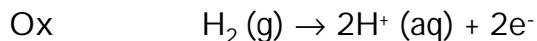
... What is it good for?

➤ Concentration Determinations



Unknown $[\text{H}^+]$
(Anode)

Std Hydrogen Electrode
(Cathode)



(1 M)

(x M)

concentration
cell

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Using Nernst

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - (0.0592/n)\text{Log}Q$$

$$Q = \frac{[\text{H}^+]^2_{\text{anode}}}{[\text{H}^+]^2_{\text{cathode}}} = \frac{[\text{H}^+]^2_{\text{anode}}}{1\text{ M}}$$

$$E_{\text{cell}} = 0.000 - (0.0592/2)\text{Log}[\text{H}^+]^2_{\text{anode}}$$

$$E_{\text{cell}} = \frac{-2(0.0592)}{2} \text{Log}[\text{H}^+]_{\text{anode}}$$

$$E_{\text{cell}} = -0.0592 \text{Log} [\text{H}^+]_{\text{anode}}$$

$$E_{\text{cell}} = 0.0592 \text{ pH}$$

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