ANSWER KEY.

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9.9 Cell Potential and Free Energy

AP Chemistry | ANSWER KEY

Show your work with units. $E°_{cell} = E°_{cathode} - E°_{anode}$ and $\Delta G° = -nFE°$, with F = 96,485 C/mol e^- and 1 V = 1 J/C.

1. Use the table of standard reduction potentials for this worksheet. For $Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$:

Half-reaction	E° (V)	Half-reaction	E° (V)
$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$	+1.36	$Ag^+(aq) + e^- \rightarrow Ag(s)$	+0.80
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	+0.34	$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$	−0.13
$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$	−0.44	$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	−0.76
$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$	−1.66	$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$	−2.37

(a) Identify the anode and cathode and calculate $E°_{cell}$. (b) Find n and calculate ΔG°.

(a) Anode: Mg (oxidized). Cathode: Cu. $E°_{cell} = E°_{cathode} - E°_{anode}$ = (+0.34 V) − (−2.37 V) = **+2.71 V**

(b) **n = 2 mol e^-.** $\Delta G° = -nFE° = -(2 \text{ mol } e^-)(96{,}485 \text{ C/mol } e^-)(2.71 \text{ J/C}) = -522{,}949 \text{ J}$ = **−523 kJ/mol$_{rxn}$**

2. Is $2Ag(s) + Zn^{2+}(aq) \rightarrow 2Ag^+(aq) + Zn(s)$ thermodynamically favored under standard conditions? Justify with $E°_{cell}$ and the sign of ΔG°.

Ag is oxidized (anode), Zn^{2+} is reduced (cathode). $E°_{cell} = (-0.76) - (+0.80)$ = **−1.56 V**

Not favored. $E°_{cell} < 0$, so $\Delta G° = -nFE° > 0$. The reverse reaction is favored, with $E°_{cell}$ = +1.56 V.

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3. For $Fe(s) + 2Ag^{+}(aq) \rightarrow Fe^{2+}(aq) + 2Ag(s)$, a student writes $E°_{cell} = 2(+0.80) - (-0.44) = +2.04$ V. Find the error and give the correct $E°_{cell}$.

E° is never multiplied by a coefficient. Potential (J/C) is an intensive property; it does not depend on how many electrons transfer.

$E°_{cell} = (+0.80) - (-0.44) =$ **+1.24 V**

4. For $2Al(s) + 3Cu^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Cu(s)$, find $E°_{cell}$, the number of electrons transferred, n, and ΔG°.

$E°_{cell} = (+0.34) - (-1.66) =$ **+2.00 V**

$2Al \rightarrow 2Al^{3+} + 6e^{-}$ and $3Cu^{2+} + 6e^{-} \rightarrow 3Cu$, so **n = 6 mol e^{-}**

$\Delta G° = -(6)(96{,}485)(2.00) = -1{,}157{,}820$ J = **-1.16×10^{3} kJ/mol_{rxn}**

5. Use the table. (a) Which species is the strongest oxidizing agent? The strongest reducing agent? (b) Will Cu(s) react with $Fe^{2+}(aq)$ under standard conditions? Justify with $E°_{cell}$.

(a) **Strongest oxidizing agent: $Cl_2(g)$** (highest E°, easiest to reduce). **Strongest reducing agent: Mg(s)** (lowest E°, easiest to oxidize).

(b) **No.** Cu would be the anode and Fe^{2+} the cathode: $E°_{cell} = (-0.44) - (+0.34) = -0.78$ V < 0, so the reaction is not favored.

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