AP® Chemistry
2002 Sample Student Responses

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2A,

Answer EITHER Question 2 below OR Question 3 printed on page 12. Only one of these two questions will be graded. If you start both questions, be sure to cross out the question you do not want graded. The Section II score weighting for the question you choose is 20 percent.

2. Answer parts (a) through (e) below, which relate to reactions involving silver ion, Ag^+.

The reaction between silver ion and solid zinc is represented by the following equation.

$$2\ Ag^+(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + 2\ Ag(s)$$

(a) A 1.50 g sample of Zn is combined with 250. mL of 0.110 M $AgNO_3$ at 25°C.

(i) Identify the limiting reactant. Show calculations to support your answer.

(ii) On the basis of the limiting reactant that you identified in part (i), determine the value of $[Zn^{2+}]$ after the reaction is complete. Assume that volume change is negligible.

(b) Determine the value of the standard potential, $E°$, for a galvanic cell based on the reaction between $AgNO_3(aq)$ and solid Zn at 25°C.

Another galvanic cell is based on the reaction between $Ag^+(aq)$ and $Cu(s)$, represented by the equation below. At 25°C, the standard potential, $E°$, for the cell is 0.46 V.

$$2\ Ag^+(aq) + Cu(s) \rightarrow Cu^{2+}(aq) + 2\ Ag(s)$$

(c) Determine the value of the standard free-energy change, $\Delta G°$, for the reaction between $Ag^+(aq)$ and $Cu(s)$ at 25°C.

(d) The cell is constructed so that $[Cu^{2+}]$ is 0.045 M and $[Ag^+]$ is 0.010 M. Calculate the value of the potential, E, for the cell.

(e) Under the conditions specified in part (d), is the reaction in the cell spontaneous? Justify your answer.

a) (i) Molar mass of Zn = 65.4 g/mol

∴ mols Zn = 1.50g ÷ 65.4 g/mol = 0.0229 mols Zn

mols Ag^+ = M x V = 0.250 L x 0.110 M = 0.275 mols Ag^+

The mol ratio of Ag^+ to Zn is 2 : 1

∴ if we started with 0.0229 mols Zn, we would require

$$0.0229 \text{ mols Zn} \times \frac{2 \text{ mol } Ag^+}{1 \text{ mol Zn}} = 0.0458 \text{ mols } Ag^+$$

Since we require 0.0458 mols Ag^+ and we only have 0.275 mols Ag^+, Ag^+ is the limiting reactant

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(ii) mols Zn^{2+} = $0.275 \text{ mols } Ag^+ \times \frac{1 \text{ mol } Zn^{2+}}{2 \text{ mol } Ag^+} = 0.138 \text{ mols } Zn^{2+}$

Since we have 250 mL of solution, and the volume change upon addition of Zn is negligible $[Zn^{2+}] = \frac{M}{V} = \frac{0.138 \text{ mols}}{0.250 L} = 0.552 M$

The concentration of Zn^{2+} is 0.552 M

b) $E° = E_{red}(\text{cathode}) - E_{red}(\text{anode})$

Half reactions:

At cathode: $Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$ $\quad E°_{red} = +0.80 V$

At anode: $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$ $\quad E°_{red} = -0.76 V$

$\therefore E° = +0.80 V - (-0.76 V) = 1.56 V$

c) $\Delta G° = -nFE°$ where n = 2 mols of electrons

$\Delta G° = -(2)(96,500)(0.46 V)$ $\quad$ F = 96,500 coulombs

$\Delta G° = -8.9 \times 10^4 J$ $\quad$ $E° = 0.46 V$

d) We use $E = E°_{cell} - \frac{RT}{nF} \ln Q$ but at 25°C $E = E°_{cell} - \frac{0.0592}{n} \log Q$

In this case: $Q = \frac{[Cu^{2+}]}{[Ag^+]^2} = \frac{(0.045 M)}{(0.010 M)^2} = 4.5 \times 10^2$

$\therefore E = (0.46 V) - \frac{0.0592}{2} \log(4.5 \times 10^2)$

$E = 0.38 V$

e) Yes, the reaction in the cell is spontaneous since E > 0. For a reaction to be spontaneous E must be greater than zero.

To check whether ΔG is negative under these conditions: to prove that the reaction is spontaneous

$\Delta G = \Delta G° + RT \ln Q$

$\Delta G = (-8.9 \times 10^4 J) + (8.31 J/mol\text{-}K)(298 K) \ln(4.5 \times 10^2)$

$\Rightarrow \Delta G < 0 \quad \therefore$ the reaction is spontaneous

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2B 1

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(d) The cell is constructed so that $[Cu^{2+}]$ is 0.045 *M* and $[Ag^+]$ is 0.010 *M*. Calculate the value of the potential, E, for the cell.

(e) Under the conditions specified in part (d), is the reaction in the cell spontaneous? Justify your answer.

a) i) $\left(1.50\,g\right)\left(\frac{1\,mol}{65\,g}\right) = .023$ mol Zn

$\left(.250\,L\right)\left(\frac{.110\,mol}{L}\right) = .028$ mol $AgNO_3$

$AgNO_3$ is the limiting reactant since in the equation Ag^+ should have twice the moles of Zn to react. Since $AgNO_3$ could only give Ag^+ less than twice the moles of Zn, all of the .028 Ag^+ moles are used while only .014 Zn moles can be used.

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ii) $(.028 \text{ mol } Ag^+)\left(\frac{1 \text{ mol } Zn^{2+}}{2 \text{ mol } Ag^+}\right) = .014 \text{ mole } Zn^{2+}$

$.014 \text{ mol} / .250 \text{ L} = .056 \text{ M} = [Zn^{2+}]$

b) $2Ag^+ + 2e^- \rightarrow 2\,Ag_{(s)}$ $2(.8)$ volts

$Zn \rightarrow Zn^{2+} + 2e^-$ $+(+.76)$ volts

$2.36V = \mathcal{E}^\circ$

c) $\Delta G = -nF\mathcal{E}^\circ$

$= -2\,(96500)(.46\,V)$

$\Delta G = -88,780$

d) $\mathcal{E}cell = \mathcal{E}^\circ - \frac{.0592}{n} \log Q$

$\mathcal{E}cell = .46 - \frac{.0592}{2} \log \frac{(.045)}{(.01)^2}$

$\mathcal{E}cell = .46 - .0785$

$\boxed{\mathcal{E}cell = .38 \text{ volts}}$

e) Yes it is spontaneous since the $\mathcal{E}$ of the cell is a (+) number. This means the potential is ~~[illegible]~~ increases - which happens when a reaction is spontaneous and needs no charge or energy to start the reaction.

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2C 1

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65.39

Another galvanic cell is based on the reaction between $Ag^+(aq)$ and $Cu(s)$, represented by the equation below. At 25°C, the standard potential, $E°$, for the cell is 0.46 V.

$$2\,Ag^+(aq) + Cu(s) \rightarrow Cu^{2+}(aq) + 2\,Ag(s)$$

(c) Determine the value of the standard free-energy change, $\Delta G°$, for the reaction between $Ag^+(aq)$ and $Cu(s)$ at 25°C.

(d) The cell is constructed so that $[Cu^{2+}]$ is 0.045 M and $[Ag^+]$ is 0.010 M. Calculate the value of the potential, E, for the cell.

(e) Under the conditions specified in part (d), is the reaction in the cell spontaneous? Justify your answer.

a) i) $2Ag^+ + Zn(s) \rightarrow Zn^{+2} + 2Ag(s)$

250 mL   1.50 g

.110 M

$1.50\,g\ Zn \times \frac{1\,mol}{65.39\,g\,Zn} = .0229\ mol\ Zn \leftarrow LR$

$250\,mL\,Ag^+ \times \frac{.110\,mol}{1000\,mL} = .0275\ mol\ Ag$ less mol, used up faster.

ii) $.0229\,mol\ Zn \times \frac{1\,mol\,Zn^{+2}}{1\,mol\,Zn} = \frac{.0229\,mol\ Zn^{+2}}{.250\,L} = .0916$

$[Zn^{+2}] = .0916M$

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b). $Ag^{+} + e^{-} \rightarrow Ag_{(s)}$ E°_{red} .80

$Zn^{+2} + 2e^{-} \rightarrow Zn_{(s)}$ E°_{red} −.76

$Zn_{(s)} \rightarrow Zn^{+2} + 2e^{-}$ E°_{oxi} .76

E°_{cell} .80

.76

1.56V

c) $\Delta G^{\circ} = -n\mathcal{F}\mathcal{E}^{\circ}$

$= (-3)(96500\ \text{coul})(.46V)$

$= -133170$

d) $E_{cell} = E^{\circ}_{cell} - \frac{0.0592}{n}\log Q$ @ 25°C

$.46V - \frac{.0592}{2}\log\frac{[.045M]}{[.010M]}$

$E_{cell} = .44$

e) yes, it is spontaneous because the Ecell is positive. This means the it will undergo w/ the reaction.

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