



## AP<sup>®</sup> Chemistry 2002 Sample Student Responses

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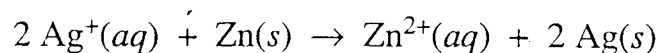
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2A,

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,  $\text{Ag}^+$ .

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



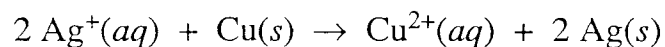
(a) A 1.50 g sample of Zn is combined with 250. mL of 0.110 M  $\text{AgNO}_3$  at 25°C.

(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  $[\text{Zn}^{2+}]$  after the reaction is complete. Assume that volume change is negligible.

(b) Determine the value of the standard potential,  $E^\circ$ , for a galvanic cell based on the reaction between  $\text{AgNO}_3(aq)$  and solid Zn at 25°C.

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



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

(d) The cell is constructed so that  $[\text{Cu}^{2+}]$  is 0.045 M and  $[\text{Ag}^+]$  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) ① Molar mass of Zn = 65.4 g/mol

$$\therefore \text{mols Zn} = 1.50 \text{ g} \div 65.4 \text{ g/mol} = 0.0229 \text{ mols Zn}$$

$$\text{mols Ag}^+ = M \times V = 0.250 \text{ L} \times 0.110 \text{ M} = 0.275 \text{ mols Ag}^+$$

The mol ratio of  $\text{Ag}^+$  to Zn is 2:1

$\therefore$  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  $\text{Ag}^+$  and we only have 0.275 mols  $\text{Ag}^+$ ,  $\text{Ag}^+$  is the limiting reactant

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2A<sub>2</sub>

## ADDITIONAL PAGE FOR ANSWERING QUESTION 2.

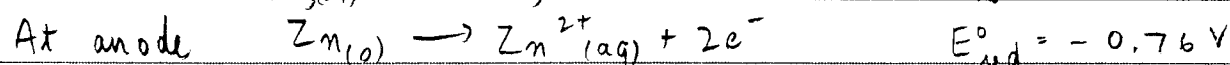
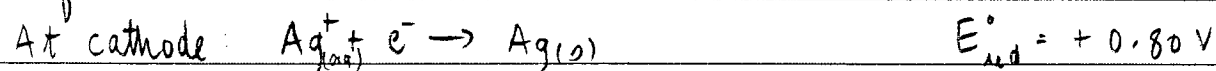
$$\textcircled{11} \quad \text{mols Zn}^{2+} = 0.275 \text{ mols Ag}^+ \times \frac{1 \text{ mol Zn}^{2+}}{2 \text{ mol Ag}^+} = 0.138 \text{ mols Zn}^{2+}$$

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

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

$$b) \quad E^\circ = E_{\text{red}}^\circ(\text{cathode}) - E_{\text{red}}^\circ(\text{anode})$$

Half reactions:



$$\therefore E^\circ = +0.80 \text{ V} - (-0.76 \text{ V}) = 1.56 \text{ V}$$

$$c) \quad \Delta G^\circ = -nFE^\circ \quad \text{where } n = 2 \text{ mols of electrons}$$

$$\Delta G^\circ = -(2)(96,500)(0.46 \text{ V}) \quad F = 96,500 \text{ coulombs}$$

$$\Delta G^\circ = -8.9 \times 10^4 \text{ J} \quad E^\circ = 0.46 \text{ V}$$

$$d) \quad \text{We use } E = E_{\text{cell}}^\circ - \frac{RT}{nF} \ln Q \quad \text{but at } 25^\circ\text{C: } E = E_{\text{cell}}^\circ - \frac{0.0592}{n} \log Q$$

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

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

$$E = 0.38 \text{ V}$$

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  $\Delta G$  is negative under these conditions: to prove that the reaction is spontaneous

$$\Delta G = \Delta G^\circ + RT \ln Q$$

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

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

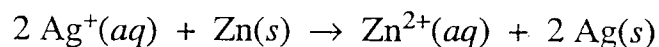
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2B1

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.

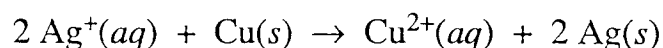
2. Answer parts (a) through (e) below, which relate to reactions involving silver ion,  $\text{Ag}^+$ .

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



- (a) A 1.50 g sample of Zn is combined with 250. mL of 0.110 M  $\text{AgNO}_3$  at 25°C.
- Identify the limiting reactant. Show calculations to support your answer.
  - On the basis of the limiting reactant that you identified in part (i), determine the value of  $[\text{Zn}^{2+}]$  after the reaction is complete. Assume that volume change is negligible.
- (b) Determine the value of the standard potential,  $E^\circ$ , for a galvanic cell based on the reaction between  $\text{AgNO}_3(aq)$  and solid Zn at 25°C.

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



- (c) Determine the value of the standard free-energy change,  $\Delta G^\circ$ , for the reaction between  $\text{Ag}^+(aq)$  and  $\text{Cu}(s)$  at 25°C.
- (d) The cell is constructed so that  $[\text{Cu}^{2+}]$  is 0.045 M and  $[\text{Ag}^+]$  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.

$$a) \frac{(1.50 \text{ g})}{(65 \text{ g})} \left( \frac{1 \text{ mol}}{65 \text{ g}} \right) = 0.023 \text{ mol Zn}$$

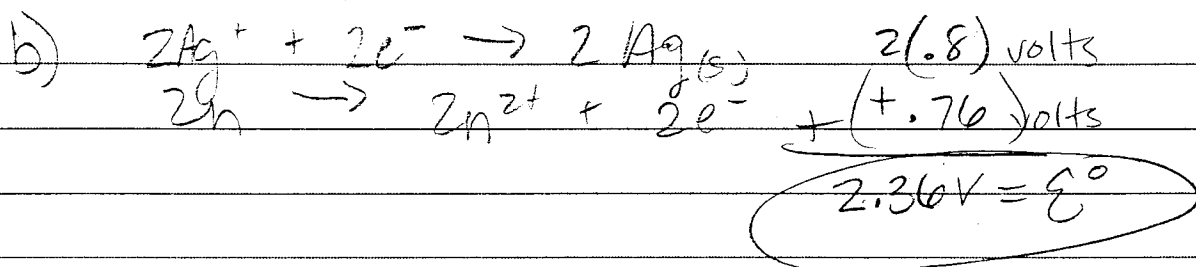
$$(0.250 \text{ L}) \left( \frac{0.110 \text{ mol}}{\text{L}} \right) = 0.028 \text{ mol AgNO}_3$$

$\text{AgNO}_3$  is the limiting reactant since ~~in the~~ in the equation  $\text{Ag}^+$  should have twice the moles of Zn to react. Since  $\text{AgNO}_3$  could only give  $\text{Ag}^+$  less than twice the moles of Zn, all of the 0.028  $\text{Ag}^+$  moles are used while only 0.014 Zn moles can be used.

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

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



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

$$= -2(96,500)(.46 \text{ V})$$

$$\Delta G = -88,780$$

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

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

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

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

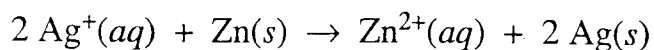
e) Yes it is spontaneous since the  $\mathcal{E}$  of the cell is a (+) number. This means the potential is ~~increases~~ increases - which happens when a reaction is spontaneous and needs no change or energy to start the reaction.

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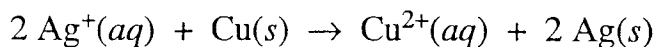
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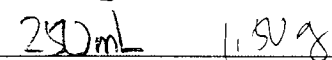
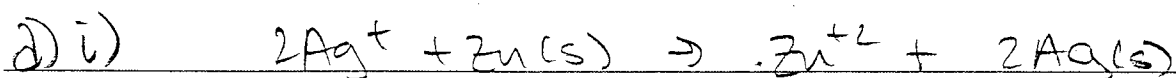


- (a) A 1.50 g sample of Zn is combined with 250. mL of 0.110 M  $\text{AgNO}_3$  at 25°C.
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- (d) The cell is constructed so that  $[\text{Cu}^{2+}]$  is 0.045 M and  $[\text{Ag}^+]$  is 0.010 M. Calculate the value of the potential,  $E$ , for the cell.
- (e) Under the conditions specified in part (d), is the reaction in the cell spontaneous? Justify your answer.



$$1.50 \text{ g Zn} \times \frac{1 \text{ mol}}{65.39 \text{ g Zn}} = 0.0229 \text{ mol Zn} \leftarrow \text{LR}$$

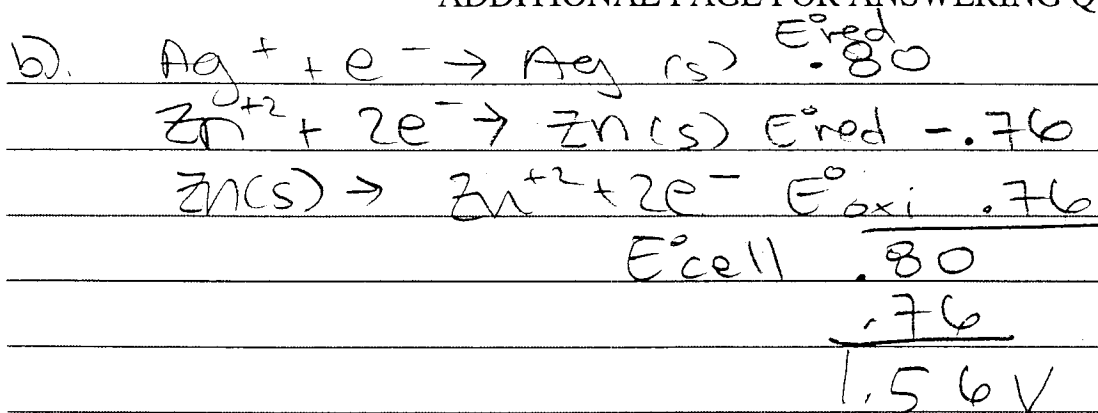
$$250 \text{ mL Ag}^+ \times \frac{0.110 \text{ mol}}{1000 \text{ mL}} = 0.0275 \text{ mol Ag} \quad \text{less mol, used up faster.}$$

ii)  $0.0229 \text{ mol Zn} \times \frac{1 \text{ mol Zn}^{2+}}{1 \text{ mol Zn}} = 0.0229 \text{ mol Zn}^{2+}$

$$\frac{0.0229 \text{ mol Zn}^{2+}}{0.250 \text{ L}} = 0.0916 \quad [\text{Zn}^{2+}] = 0.0916 \text{ M}$$

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2C2



$$\begin{aligned}
 \text{c) } \Delta G^{\circ} &= -n F E^{\circ} \\
 &= (-3)(96500 \text{ cal})(.46 \text{ V}) \\
 &= -133170
 \end{aligned}$$

$$\begin{aligned}
 \text{d) } E_{\text{cell}} &= E^{\circ}_{\text{cell}} - \frac{0.0592}{n} \log Q \quad @ 25^{\circ}\text{C} \\
 & .46 \text{ V} - \frac{.0592}{2} \log \frac{[.045 \text{ M}]}{[.010 \text{ M}]}
 \end{aligned}$$

$$E_{\text{cell}} = .44$$

e) yes, it is spontaneous because the  $E_{\text{cell}}$  is positive. This means the it will undergo w/ the reaction.

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