AP Chemistry: Electrochemistry
Multiple Choice Answers

14. Questions 14-17
The spontaneous reaction that occurs when the cell in the picture operates is as follows:

$$2\text{Ag}^+ + \text{Cd(s)} \rightarrow 2\text{Ag(s)} + \text{Cd}^{2+}$$

(A) Voltage increases.
(B) Voltage decreases but remains > zero.
(C) Voltage becomes zero and remains at zero.
(D) No change in voltage occurs.
(E) Direction of voltage change cannot be predicted without additional information.

Which of the above occurs for each of the following circumstances?

14. A 50-milliliter sample of a 2-molar \(\text{Cd(NO}_3\text{)}_2\) solution is added to the left beaker. **B – voltage decreases but remains above zero**, voltage decreases because of the increase in the concentration values

15. The silver electrode is made larger. **D – No change because the size of the electrode is not provided in the initial problem**

16. The salt bridge is replaced by a platinum wire. **C – removing the salt bridge stops the flow of electrons resulting in a voltage of zero, and electrons cannot flow through a platinum wire so, the voltage will remain zero**

17. Current is allowed to flow for 5 minutes. **B – voltage decreases but remains above zero**

29. \(\text{Cu}_\text{(s)} + 2\text{Ag}^+ \rightarrow \text{Cu}^{2+} + 2\text{Ag}_\text{(s)}\)
If the equilibrium constant for the reaction above is \(3.7 \times 10^{15}\), which of the following correctly describes the standard voltage, \(E^\circ\), and the standard free energy change, \(\Delta G^\circ\), for this reaction?

(A) \(E^\circ\) is positive and \(\Delta G^\circ\) is negative.
(B) \(E^\circ\) is negative and \(\Delta G^\circ\) is positive.
(C) \(E^\circ\) and \(\Delta G^\circ\) are both positive.
(D) \(E^\circ\) and \(\Delta G^\circ\) are both negative.
(E) \(E^\circ\) and \(\Delta G^\circ\) are both zero

>> If the equilibrium constant is greater than 1, then the log of that number will be positive resulting in a \(-\Delta G\) and a positive \(E^\circ\)

46. If 0.060 faradays are passed through an electrolytic cell containing a solution of \(\text{In}^{3+}\) ions, the maximum number of moles of In that could be deposited at the cathode is...

(A) 0.010 mole  (B) **0.020 mole**  (C) 0.030 mole  (D) 0.060 mole  (E) 0.18 mole

>> A faraday = 1 mole of \(e^-\)

\(\text{In}^{3+} + 3e^- \rightarrow \text{In}\)

> 3 moles of \(e^-\)

\(> 0.060 \text{ mol } e^- \times (1 \text{ mol In/3mol } e^-) = 0.020 \text{ moles In}\)
75. If a copper sample containing some zinc impurity is to be purified by electrolysis, the anode and the cathode must be which of the following?

<table>
<thead>
<tr>
<th>Anode</th>
<th>Cathode</th>
</tr>
</thead>
<tbody>
<tr>
<td>(A) Pure copper</td>
<td>Pure zinc</td>
</tr>
<tr>
<td>(B) Pure zinc</td>
<td>Pure copper</td>
</tr>
<tr>
<td>(C) Pure copper</td>
<td>Impure copper sample</td>
</tr>
<tr>
<td>(D) Impure copper sample</td>
<td>Pure copper</td>
</tr>
<tr>
<td>(E) Impure copper sample</td>
<td>Pure zinc</td>
</tr>
</tbody>
</table>

Fe\(^{2+}\) + 2e\(^{-}\) → Fe\(_{(s)}\) \(\text{E}^\circ = -0.44 \text{ volt}\)

Ni\(^{2+}\) + 2e\(^{-}\) → Ni\(_{(s)}\) \(\text{E}^\circ = -0.23 \text{ volt}\)

60. The standard reduction potentials for two half reactions are given above. The Nernst equation for a galvanic cell at 25° C in which Fe\(_{(s)}\) reduces Ni\(^{2+}\) is the following:

\[ E = E^\circ - \frac{0.059}{2} \log \frac{\text{[Fe}^{2+}\text{]}}{\text{[Ni}^{2+}\text{]}} \]

**What is the equilibrium constant for the reaction below?**

Fe\(_{(s)}\) + Ni\(^{2+}\) → Fe\(^{2+}\) + Ni\(_{(s)}\)

(A) \(1.9 \times 10^{-23}\)  (B) \(7.6 \times 10^{-8}\)  (C) \(3.6 \times 10^{3}\)  (D) \(1.3 \times 10^{7}\)  (E) \(5.2 \times 10^{-22}\)

>> At equilibrium, \(E = 0 \text{ V}\)

>> Anode: Fe → Fe\(^{2+}\) + 2e\(^{-}\), \(E = +0.44 \text{ V}\)
>> Cathode: Ni\(^{2+}\) + 2e\(^{-}\) → Ni, \(E = -0.23 \text{ V}\)
>> \(E^\circ = E_{\text{red}} + E_{\text{ox}} = -0.23 \text{ V} + 0.44 \text{ V} = 0.21 \text{ V}\)

>> \(0 = 0.21 - 0.0592/2 \log \frac{\text{[Fe}^{2+}\text{]}}{\text{[Ni}^{2+}\text{]}}\)
>> \(0 = 0.21 - 0.03 \log \frac{\text{[Fe}^{2+}\text{]}}{\text{[Ni}^{2+}\text{]}}\)
>> \(-0.21 = -0.03 \log \frac{\text{[Fe}^{2+}\text{]}}{\text{[Ni}^{2+}\text{]}}\)
>> \(-0.21/-0.03 = \log \frac{\text{[Fe}^{2+}\text{]}}{\text{[Ni}^{2+}\text{]}}\)
>> \(7 = \log \frac{\text{[Fe}^{2+}\text{]}}{\text{[Ni}^{2+}\text{]}}\)
>> \([\text{Fe}^{2+}]\text{/[Ni}^{2+}] = 10^{7}\)
>> Select answer D, \(1.3 \times 10^{7}\)
36. \( \text{Zn(s)} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu(s)} \)

An electrolytic cell based on the reaction represented above was constructed from zinc and copper half-cells. The observed voltage was found to be 1.00 volt instead of the standard cell potential, \( E^\circ \), of 1.10 volts. Which of the following could correctly account for this observation?

(A) The copper electrode was larger than the zinc electrode.
(B) The \( \text{Zn}^{2+} \) electrolyte was \( \text{Zn(NO}_3\text{)}_2 \), while the \( \text{Cu}^{2+} \) electrolyte was \( \text{CuSO}_4 \).
(C) The \( \text{Zn}^{2+} \) solution was more concentrated than the \( \text{Cu}^{2+} \) solution.
(D) The solutions in the half-cells had different volumes.
(E) The salt bridge contained \( \text{KCl} \) as the electrolyte.

63. Which of the following expressions is correct for the maximum mass of copper, in grams, that could be plated out by electrolyzing aqueous \( \text{CuCl}_2 \) for 16 hours at a constant current of 3.0 amperes? (1 faraday = 96,500 coulombs)

(A) \( [(16)(3,600)(3.0)(63.55)(2)] / (96,500) \)

(B) \( [(16)(3,600)(3.0)(63.55)] / [(96,500)(2)] \)

(C) \( [(16)(3,600)(3.0)(63.55)] / (96,500) \)

(D) \( [(16)(60)(3.0)(96,500)(2)] / (63.55) \)

(E) \( [(16)(60)(3.0)(96,500)] / [(63.55)(2)] \)

75. A direct-current power supply of low voltage (less than 10 volts) has lost the markings that indicate which output terminal is positive and which is negative. A chemist suggests that the power supply terminals be connected to a pair of platinum electrodes that dip into 0.1-molar \( \text{KI} \) solution. Which of the following correctly identifies the polarities of the power supply terminals?

(A) A gas will be evolved only at the positive electrode.
(B) A gas will be evolved only at the negative electrode.
(C) A brown color will appear in the solution near the negative electrode.
(D) A metal will be deposited on the positive electrode.
(E) None of the methods above will identify the polarities of the power supply terminals.

\[ \text{Cu}^{2+} + 2e^- \rightarrow \text{Cu} \]

\[ \text{3 amps} = 3 \text{ C/sec} \]

\[ \text{16 hrs} \times (3600 \text{ sec/hr}) \times (3 \text{ C/sec}) \times (1 \text{ mol e}^- / 96500 \text{ C}) \times (1 \text{ mol Cu} / 2 \text{ mol e}^-) \times (63.55 \text{ g Cu} / \text{ mol Cu}) = [16 \times 3600 \times 3 \times 63.55] / [96500 \times 2] \]
Questions 34-35 refer to an electrolytic cell that involves the following half-reaction:
\[ \text{AlF}_6^{3-} + 3 \text{e}^- \rightarrow \text{Al} + 6\text{F}^- \]

34. Which of the following occurs in the reaction?
(A) \( \text{AlF}_6^{3-} \) is reduced at the cathode.
(B) Al is oxidized at the anode.
(C) Aluminum is converted from the \(-3\) oxidation state to the 0 oxidation state.
(D) \( \text{F}^- \) acts as a reducing agent.
(E) \( \text{F}^- \) is reduced at the cathode.

35. As steady current of 10 amperes in passed though an aluminum-production cell for 15 minutes. Which of the following is the correct expression for calculating the number of grams of aluminum produced? (1 faraday = 96,500 coulombs)

\[
\begin{align*}
\text{Al}^{3+} + 3\text{e}^- & \rightarrow \text{Al} \\
\text{3 mole e}^- & \\
\text{10 amps} = 10 \text{ C / sec} \\
\text{15 min} \times (60 \text{ sec/min}) \times (10 \text{ C / sec}) \times (1 \text{ mol e}^- / 96500 \text{ C}) \times (1 \text{ mol Al} / 3 \text{ mol e}^-) \times (27 \text{ g Al} / 1 \text{ mol Al}) \\
\text{Answer C}
\end{align*}
\]

\[
\begin{align*}
\text{M}_{(s)} + 3 \text{Ag}^{+}_{(aq)} & \rightarrow 3 \text{Ag}_{(s)} + \text{M}^{3+}_{(aq)} & \text{E} = +2.46 \text{ V} \\
\text{Ag}^{+}_{(aq)} + \text{e}^- & \rightarrow \text{Ag}_{(s)} & \text{E} = +0.80 \text{ V}
\end{align*}
\]

57. According to the information above, what is the standard reduction potential for the half-reaction \( \text{M}^{3+}_{(aq)} + 3 \text{e}^- \rightarrow \text{M}_{(s)} \)?
(A) \(-1.66 \text{ V}\) (B) \(-0.06 \text{ V}\) (C) \(0.06 \text{ V}\) (D) \(1.66 \text{ V}\) (E) \(3.26 \text{ V}\)

\[
\begin{align*}
\text{E}_{\text{cell}} & = \text{E}_{\text{red}} + \text{E}_{\text{ox}} \\
\text{2.46 V} & = 0.80 + \text{E}_{\text{ox}} \\
\text{E}_{\text{ox}} & = +1.66 \text{ V} \text{ but need to reverse the sign to make the potential be a reduction potential so the answer would be } -1.66 \text{ V}
\end{align*}
\]
20. \[ \text{Mg(s)} + \text{NO}_3^-(aq) + \text{H}^+(aq) \rightarrow \text{Mg}^{2+}(aq) + \text{NH}_4^+(aq) + \text{H}_2\text{O(l)} \]

When the skeleton equation above is balanced and all coefficients reduced to their lowest whole-number terms, what is the coefficient for H\(^+\)?

(A) 4 \hspace{1cm} (B) 6 \hspace{1cm} (C) 8 \hspace{1cm} (D) 9 \hspace{1cm} (E) 10

\[ \text{Red: NO}_3^-(aq) + 10 \text{H}^+(aq) + 8 \text{e}^- \rightarrow \text{NH}_4^+(aq) + 3 \text{H}_2\text{O(l)} \]

\[ \text{Ox: } 4 \text{Mg(s)} \rightarrow 4 \text{Mg}^{2+}(aq) + 2 \text{e}^- \]

\[ \text{SUM: NO}_3^-(aq) + 10 \text{H}^+(aq) + 4 \text{Mg(s)} \rightarrow 4 \text{Mg}^{2+} + \text{NH}_4^+(aq) + 3 \text{H}_2\text{O(l)} \]

34. \[ \text{CrO}_2^{2-} + \text{OH}^- \rightarrow \text{CrO}_4^{2-} + \text{H}_2\text{O} + \text{e}^- \]

When the equation for the half-reaction above is balanced, what is the ratio of the coefficients \(\text{OH}^- / \text{CrO}_2^{2-}\)?

(A) 1:1 \hspace{1cm} (B) 2:1 \hspace{1cm} (C) 3:1 \hspace{1cm} (D) 4:1 \hspace{1cm} (E) 5:1

\[ \text{CrO}_2^{2-} \rightarrow \text{CrO}_4^{2-} \text{(skeleton without water or OH}^-\text{)} \]

\[ \text{CrO}_2^{2-} + 2 \text{H}_2\text{O} \rightarrow \text{CrO}_4^{2-} + 4 \text{H}^+ + 3 \text{e}^- + 4 \text{OH}^- + \text{OH}^- \text{(makes 4 waters on the right, cancels 2 waters on the left)} \]

\[ \text{SUM: CrO}_2^{2-} + 4 \text{OH}^- \rightarrow \text{CrO}_4^{2-} + 2 \text{H}_2\text{O} + 3 \text{e}^- \]

61. When a solution of potassium dichromate is added to an acidified solution of iron(II) sulfate, the products of the reaction are...

(A) \text{FeCr}_2\text{O}_7(s) and \text{H}_2\text{O} \hspace{1cm} (B) \text{FeCrO}_4(s) and \text{H}_2\text{O} \hspace{1cm} (C) \text{Fe}^{3+}, \text{CrO}_4^{2-}, \text{and H}_2\text{O}

(D) \text{Fe}^{3+}, \text{Cr}^{3+}, \text{and H}_2\text{O} \hspace{1cm} (E) \text{Fe}_2(\text{SO}_4)_3(s), \text{Cr}^{3+} \text{and H}_2\text{O}

\[ \text{It is a classic reaction that dichromate will be reduced to Cr}^{3+} \text{ in acid solution. That leaves iron to be oxidized to from Fe}^{2+} \text{ to Fe}^{3+} \text{ and if H}^+ \text{ is present on one side of a redox reaction, water is present on the other} \]

65. Which of the following species CANNOT function as an oxidizing agent?

(A) \text{Cr}_2\text{O}_7^{2-} \hspace{1cm} (B) \text{MnO}_4^- \hspace{1cm} (C) \text{NO}_3^- \hspace{1cm} (D) \text{S} \hspace{1cm} (E) \text{I}^-\]

\[ \text{An oxidizing agent is itself reduced. Therefore, you are really being asked, “Which species cannot be reduced?” I}^- \text{ cannot be reduced, all of its other oxidation states are either zero or positive, which requires oxidation.} \]

79. \[ 5 \text{Fe}^{2+} + \text{MnO}_4^- + 8 \text{H}^+ \rightleftharpoons 5 \text{Fe}^{3+} + \text{Mn}^{2+} + 4 \text{H}_2\text{O} \]

In a titration experiment based on the equation above, 25.0 milliliters of an acidified \text{Fe}^{2+} \text{ solution requires 14.0 milliliters of standard 0.050-molar MnO}_4^- \text{ solution to reach the equivalence point. The concentration of Fe}^{2+} \text{ in the original solution is...}

(A) 0.0035 M \hspace{1cm} (B) 0.0070 M \hspace{1cm} (C) 0.035 M \hspace{1cm} (D) 0.070 M \hspace{1cm} (E) 0.14 M

\[ 5 \text{Fe}^{2+} + \text{MnO}_4^- + 8 \text{H}^+ \rightleftharpoons 5 \text{Fe}^{3+} + \text{Mn}^{2+} + 4 \text{H}_2\text{O} \]

\[ \text{Calculate the number of moles, } (.05 \text{ mol/L x 14 mL}) = 0.7 \text{ mmol of MnO}_4^- \]

\[ \text{So, 5 times that amount is the required mmol of Fe}^{2+}. \text{ That’s 3.5 mmol, so the molarity is 3.5 mmol/100 mL (the “milli” cancels) = 0.035 M.} \]
20. \[ 6 \text{I}^- + 2 \text{MnO}_4^- + 4 \text{H}_2\text{O(l)} \rightarrow 3 \text{I}_2(s) + 2 \text{MnO}_2(s) + \text{OH}^- \]

Which of the following statements regarding the reaction represented by the equation above is correct?
(A) Iodide ion is oxidized by hydroxide ion.
(B) MnO$_4^-$ is oxidized by iodide ion.
(C) The oxidation number of manganese changes from +7 to +2.
(D) The oxidation number of manganese remains the same.
(E) The oxidation number of iodine changes from −1 to 0.

22. \[ \text{Cr}_2\text{O}_7^{2-} + 2\text{e}^- + 14\text{H}^+ \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O(l)} \]

When the equation for the half reaction above is balanced with the lowest whole-number coefficients, the coefficient for H$_2$O is...
(A) 2  (B) 4  (C) 6  (D) 7  (E) 14

61. When acidified K$_2$Cr$_2$O$_7$ solution is added to Na$_2$S solution, green Cr$_3^+$ ions and free S are formed. When acidified K$_2$Cr$_2$O$_7$ solution is added to NaCl, no change occurs. Of the substances involved in these reactions, which is the best reducing agent?
(A) K$_2$Cr$_2$O$_7$  (B) Na$_2$S  (C) Cr$_3^+$  (D) S  (E) NaCl

18. \[ 2\text{H}_2\text{O} + 4\text{MnO}_4^- + 3\text{ClO}_2^- \rightarrow 4\text{MnO}_2 + 3\text{ClO}_4^- + 4\text{OH}^- \]

Which species acts as an oxidizing agent in the reaction represented above?
(A) H$_2$O  (B) ClO$_4^-$  (C) ClO$_2^-$  (D) MnO$_2$  (E) MnO$_4^-$

20. \[ 6\text{Ag}^+ + \text{AsH}_3(g) + 6\text{OH}^- \rightarrow 6\text{Ag(s)} + \text{H}_3\text{AsO}_3(aq) + 3\text{H}_2\text{O} \]

When the equation above is balanced with lowest whole-number coefficients, the coefficient for OH$^-$ is...
(A) 2  (B) 4  (C) 5  (D) 6  (E) 7

29. When an aqueous solution of NaOH is added to an aqueous solution of potassium dichromate, K$_2$Cr$_2$O$_7$ the dichromate ion is converted to...
(A) CrO$_4^{2-}$  (B) CrO$_2$  (C) Cr$_3^+$  (D) Cr$_2$O$_3(s)$  (E) Cr(OH)$_3(s)$

>> In a basic solution dichromate is oxidized to the chromate ion while in acidic solution dichromate is oxidized to the Cr$_3^+$ ion

71. \[ \text{O}_2 + \text{Fe(OH)}_2 + \text{H}_2\text{O} \rightarrow \text{Fe(OH)}_3 \]

If 1 mole of O$_2$ oxidizes Fe(OH)$_2$ according to the reaction represented above, how many moles of Fe(OH)$_3$ can be formed?
(A) 2  (B) 3  (C) 4  (D) 5  (E) 6

>> \[ 4\text{Fe}^{2+} + \text{O}_2 + 2\text{H}^+ \rightarrow 4\text{Fe}^{3+} + 2\text{OH}^- \]
20. What mass of Au is produced when 0.0500 mol of $Au_2S_3$ is reduced completely with excess $H_2$?  
(A) 9.85 g  (B) 19.7 g  (C) 24.5 g  (D) 39.4 g  (E) 48.9 g

$$Au_2S_3 + H_2 \rightarrow Au + H_xS_y$$

If 0.0500 moles of $Au_2S_3$ completely reacts (not a limiting reactant), then 0.100 moles of Au were formed.

$$\text{The MM of Au is 197, so 0.100 mol} \times 197 \text{ g/mol} = 19.7 \text{ g}$$

42. $\ldots Li_3N(s) + \ldots H_2O(l) \rightarrow \ldots Li^+(aq) + \ldots OH^-(aq) + \ldots NH_3(g)$

When the equation above is balanced and all coefficients reduced to lowest whole number terms, the coefficient for $OH^-(aq)$ is...

(A) 1  (B) 2  (C) 3  (D) 4  (E) 6

$$\text{Li}_3N(s) + 3 H_2O(l) \rightarrow Li^+(aq) + 3 OH^-(aq) + NH_3(g)$$

30. Which of the following is true regarding the reaction represented above?  
(A) The oxidation number of O does not change.  
(B) The oxidation number of H changes from $-1$ to $+1$.  
(C) The oxidation number of F changes from $+1$ to $-1$.  
(D) The oxidation number of Se changes from $-2$ to $+6$.  
(E) It is a disproportionation reaction for F.

71. In the electroplating of nickel, 0.200 faraday of electrical charge is passed through a solution of $NiSO_4$. What mass of nickel is deposited?  
(A) 2.94 g  (B) 5.86 g  (C) 11.7 g  (D) 58.7 g  (E) 294 g

$$\text{Ni}^{2+} + 2 e^- \rightarrow Ni$$

faraday = 1 mol e$^{-}$

$$0.200 \text{ mol e}^- \times (1 \text{ mol Ni} / 2 \text{ mol e}^-) \times (59 \text{ g Ni} / 1 \text{ mol Ni}) = 5.9 \text{ g Ni}$$

19. In which of the following species does sulfur have the same oxidation number as it does in $H_2SO_4$?  
$$\text{In } H_2SO_4, \text{ Sulfur has an oxidation number of +6}$$

(A) $H_2SO_3$  (B) $S_2O_5^{2-}$  (C) $S^{2-}$  (D) $S_8$  (E) $SO_2Cl_2$