Galvanic (Voltaic) Cells

1988

Average score = 5.02

a) two points
   Sn ---> Sn^{2+} + $2e^-$
   Ag^+ + e^- ---> Ag
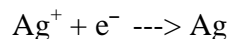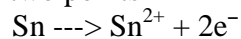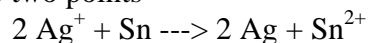
b) two points
   $2 Ag^+$ + Sn ---> 2 Ag + Sn^{2+}
   E° = [0.80 - (- 0.14)] V = 0.94 V

c) two points
   E = (0.0591 ÷ n) log K (or - nFE = - RT ln K)

   log K = (0.94 x 2) ÷ 0.0591 = 31.8

   K = 6 x 10^{31}

d) two points
   E = E° - (0.0591 ÷ n) log $[Sn^{2+}]$ / $[Ag^+]^2$
   OR
   E = E° - (RT / nF) ln Q
   Q = $[Sn^{2+}]$ / $[Ag^+]^2$
   E = 0.94 = (0.0591 ÷ 2) log (1 ÷ (0.02^2))
   E = 0.94 - 0.10 = 0.84 V
   The final point is for completion of calculations in (c) and (d).

Question 7   Standards   (8pts)

a) $2 Cr + 3 Cu^{2+} \rightarrow 2 Cr^{3+} + 3 Cu$ 1pt

Cr = reducing agent   1pt

Cu^{2+} = oxidizing agent   1pt

b) i) Cu is cathode   1pt

ii) salt bridge   1pt

iii) transfer of ions or charge but not electrons   1pt

c) Nernst equation we proposed   1pt

E decreases   1pt

Guidelines: c) LeChatlier type argument okay.

less spontaneous, less forward rxn, more reverse rxn.

If wrong rxn. written, consistency with incorrect rxn. is required. If wrong rxn is not a redox reaction, points in bi and c can only be earned if a detailed explanation accompanies. If rxn does not have both an oxidation and a reduction then no credit can be earned for agents or cathode
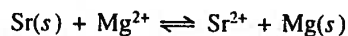
If in part a, reduction and oxidation are correctly labeled but agents are not addressed 1pt can be earned from the "agent" points.

QUESTION 7
(8 points)

7. $$Sr(s) + Mg^{2+} \rightleftharpoons Sr^{2+} + Mg(s)$$

Consider the reaction represented above that occurs at 25°C. All reactants and products are in their standard states. The value of the equilibrium constant, K_{eq}, for the reaction is 4.2×10^{17} at 25°C.

(a) Predict the sign of the standard cell potential, $E°$, for a cell based on the reaction. Explain your prediction.

(b) Identify the oxidizing agent for the spontaneous reaction.

(c) If the reaction were carried out at 60°C instead of 25°C, how would the cell potential change? Justify your answer.

(d) How would the cell potential change if the reaction were carried out at 25°C with a 1.0-molar solution of $Mg(NO_3)_2$ and a 0.10-molar solution of $Sr(NO_3)_2$? Explain.

(e) When the cell reaction in (d) reaches equilibrium, what is the cell potential?

Scoring Guide

(a) The sign of the cell potential will be positive   *1 point*

because
$\left\{ \begin{array}{l} K \text{ is greater than 1} \\ \text{the reaction is spontaneous (occurs)} \\ E^0 \text{ for } Sr^{2+} \text{ is more positive} \\ \text{Standard reduction potential for Sr more negative} \\ E^0 = +0.52 \text{ V} \end{array} \right\}$ *any one*   *1 point*

(Note: only 1 point earned for just E^0 positive because K_{eq} positive)

(b) The oxidizing agent is Mg^{2+} *1 point*

(c) The cell potential would increase   *1 point*

Since all ions are at $1\,M$, Q for the system is 1 and $E^0 = \dfrac{RT}{nF} \ln K$,

so as T increases, so should E^0 *1 point*

(Note: no credit lost if student recognizes K_{eq} dependence on T. For temperature change in this problem, decrease in $\ln K$ term is small relative to the term $\dfrac{RT}{nF}$)

OR,

No change, because in the Nernst equation $E_{cell} = E^0 - \dfrac{RT}{nF} \ln Q$,

$\ln Q = 0$, and $E_{cell} = E^0$

(Note: this approach earns 1 point only)

(d) E_{cell} will increase   *1 point*

In the equation $E_{cell} = E^0 - \dfrac{0.0592}{n} \log Q$, $Q = 0.1 \Rightarrow$ *1 point*

log Q is negative $\Rightarrow$ term after E^0 is positive $\Rightarrow$ E_{cell} increases

OR,

with the concentration of Mg^{2+} larger than that of Sr^{2+}, Le Chatelier's
principle predicts the reaction will have a larger driving force to the right   *1 point*
and a more positive E_{cell}

(e)   At equilibrium, $E_{cell} = 0$ *1 point*

(Note: "balanced", "neutral", or "no net reaction" not accepted)

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Question 8

(8 points)

(a) $2 \, Ag^+(aq) + Cd(s) \; \rightarrow \; 2 \, Ag(s) + Cd^{2+}(aq)$ *1 pt*
  • *equation must be balanced and net ionic, phases not necessary*
  • *reaction direction and ion charges must be correct*

  $0.80 - (-0.40) \; = \; 1.20 \, V$ *1 pt*
  • *evidence of where numbers came from should be present; if equation*
    is exactly reversed, -1.20 V earns the point

(b) Anions (or NO_3^- ions) will flow to the Cd^{2+} solution or from the Ag^+ solution to
  balance the charges
  OR   *1 pt*
  Anions will flow to the left to balance the positive charge of the new Cd^{2+} ions
  • *both the correct direction and justification needed to earn this point*
  • *direction may be indicated by arrow marked on diagram*

(c) The cell voltage will increase.   *1 pt*

  Ag^+ is a reactant, so increasing $[Ag^+]$ will increase the driving force (stress)
  for the forward (spontaneous) reaction and the potential will increase
  OR   *1 pt*
  Since $Q = [Cd^{2+}]/[Ag^+]^2$, increasing $[Ag^+]$ will decrease Q. According to the
  Nernst equation, $E = E° - (0.0592 \log Q)/n$, if Q decreases, then voltage increases.

(d) The cell voltage will decrease.   *1 pt*

  Adding NaCl will have no effect on the Cd cell, but will cause AgCl to precipitate
  in the Ag cell ($Ag^+ + Cl^- \rightarrow AgCl$). Thus $[Ag^+]$ decreases, and since Ag^+ is a   *1 pt*
  reactant, decreasing $[Ag^+]$ causes a decrease in voltage.
  • *Credit earned for " decreasing $[Ag^+]$ results in decreased voltage" or*
    "opposite of part (c)"

(e) Since $Q = [Cd^{2+}]/[Ag^+]^2$, diluting both solutions by the same amount will increase
  the value of Q. According to the Nernst equation, $E = E° - (0.0592 \log Q)/n$, if Q *1 pt*
  increases, then voltage decreases.
  • *No credit earned for "since the solutions are diluted, the voltage will decrease"*

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2001 SCORING GUIDELINES

Question 7

(8 points)

(a)   The anode is the electrode on the right (Zn is the anode)   ***1 point***

- Point is also earned if the student points to the Zn cell in the diagram.

The half-reaction is $Zn \rightarrow Zn^{2+} + 2\,e^-$ ***1 point***

(b) $Zn + Ni^{2+} \rightarrow Zn^{2+} + Ni$ ***1 point***

$E^{\circ}_{cell} = (-0.25\ V) - (-0.76\ V) = \mathbf{0.51\ V}$ ***1 point***

- Some work must be shown to support the answer.

(c) E_{cell} would decrease   ***1 point***

Since Ni^{2+} is a reactant, a decrease in its concentration decreases the   ***1 point***
driving force for the forward reaction

or

$$E_{cell} = E^{\circ} - \frac{RT}{n^{\wedge}} \ln Q, \text{ where } Q = \frac{[Zn^{2+}]}{[Ni^{2+}]}$$

Decreasing the $[Ni^{2+}]$ would increase the value of Q, so a larger number
would be subtracted from E°, thus decreasing the value of E_{cell}.

(d) $K > 1$ ***1 point***

E° is positive, so $K > 1$ ***1 point***

Note: The student's score in part (d) is based on the sign of E°_{cell} calculated
in part (b).

Note on Overall Question: If in part (a) a student incorrectly identifies Ni as being oxidized,
partial credit is earned if subsequent parts are followed through consistently.

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Question 7

8 points

The diagram below shows the experimental setup for a typical electrochemical cell that contains two standard half-cells. The cell operates according to the reaction represented by the following equation.

$$Zn(s) \; + \; Ni^{2+}(aq) \; \rightarrow \; Ni(s) \; + \; Zn^{2+}(aq)$$
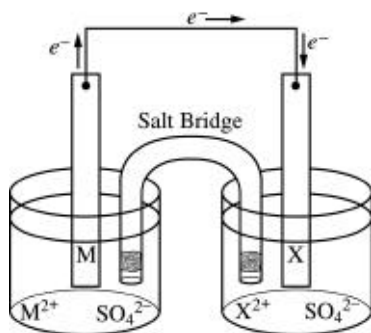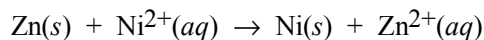
(a) Identify M and M^{2+} in the diagram and specify the initial concentration for M^{2+} in solution.

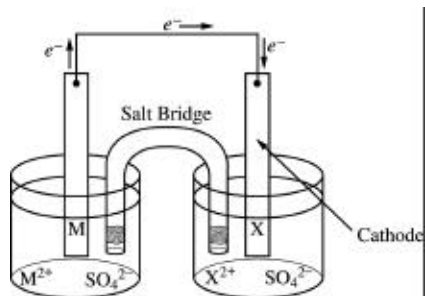|	
Electrons flow from the anode to the cathode in a voltaic electrochemical cell. The anode is where oxidation occurs, and in the reaction above, $Zn(s)$ is oxidized. So, the anode electrode must be Zn (M) and the solution contains Zn^{2+} (M^{2+}). The $[Zn^{2+}] = 1.0$ M in a standard cell. Additionally, the reduction potential for the Zn^{2+}/Zn redox couple is less than that for Ni^{2+}/Ni.	1 point earned for correct M and M^{2+} 1 point for the correct concentration of M^{2+} (Zn^{2+})

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Question 7 (cont'd.)

(b) Indicate which of the metal electrodes is the cathode. Write the balanced equation for the reaction that occurs in the half-cell containing the cathode.

The cathode is Ni(s), indicated by "X" The half-reaction is $2\,e^- + Ni^{2+} \rightarrow Ni(s)$.	1 point earned for labeling the cathode in the cell diagram 1 point earned for the half-reaction

(c) What would be the effect on the cell voltage if the concentration of Zn^{2+} was reduced to 0.100 M in the half-cell containing the Zn electrode?

$E_{cell} = E° - \dfrac{RT}{nF}\ln\dfrac{[Zn^{2+}]}{[Ni^{2+}]}$ When the $[Zn^{2+}]$ is lowered to 0.100 M, then $Q < 1$. The value of the cell potential under these nonstandard conditions is more positive than $E°$ (under standard conditions). The cell voltage increases. An argument involving LeChâtelier's principle is also acceptable: the decreased $[Zn^{2+}]$ increases the "potential" for the reaction to proceed to the right.	1 point earned for indicating that E_{cell} increases (is larger) 1 point earned for recognizing that $Q < 1$ and/or that the term $\dfrac{RT}{nF}\ln\dfrac{[Zn^{2+}]}{[Ni^{2+}]}$ must be added to the $E°$

(d) Describe what would happen to the cell voltage if the salt bridge was removed. Explain.

The cell voltage drops to zero when the salt bridge is removed. This happens because the salt bridge is needed to allow charge balance to occur in the solutions the electrodes are immersed in. In the absence of the salt bridge, ions cannot flow to balance the buildup of cations in the anode compartment and the buildup of anions in the cathode compartment.	1 point earned for the effect 1 point earned for the explanation

Question 6

Total Score 9 points

6. Answer the following questions about electrochemistry.

(a) Several different electrochemical cells can be constructed using the materials shown below. Write the balanced net-ionic equation for the reaction that occurs in the cell that would have the greatest positive value of $E°_{cell}$.

1.0 M Al(NO$_3$)$_3$ 1.0 M Cu(NO$_3$)$_2$ 1.0 M Fe(NO$_3$)$_2$

Al Metal Strip   Cu Metal Strip   Fe Metal Strip   Materials for Salt Bridge   Solution to Fill Salt Bridge   Voltmeter with Wire

$Al(s) \rightarrow Al^{3+}(aq) + 3\ e^-$ $Cu^{2+}(aq) + 2\ e^- \rightarrow Cu(s)$ $2\ Al(s) + 3\ Cu^{2+}(aq) \rightarrow 2\ Al^{3+}(aq) + 3\ Cu(s)$	1 point for selection of correct two redox couples 1 point for correctly balanced net ionic equation

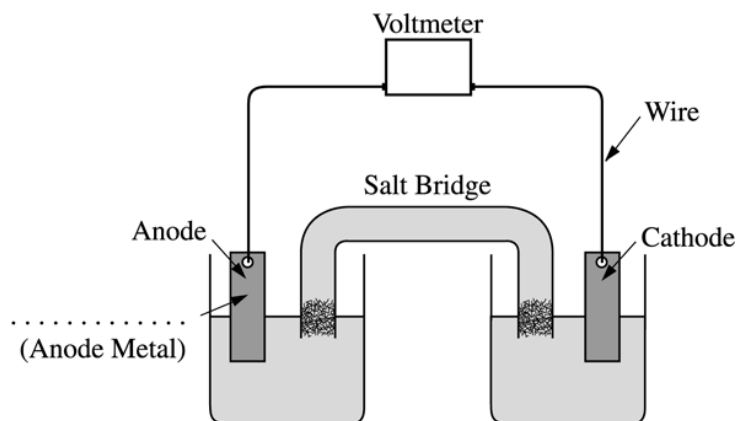(b) Calculate the standard cell potential, $E°_{cell}$, for the reaction written in part (a).

$Al^{3+}(aq) + 3\ e^- \rightarrow Al(s)$ $E° = -1.66$ V $Cu^{2+}(aq) + 2\ e^- \rightarrow Cu(s)$ $E° = +0.34$ V $E°_{cell} = E°_{cathode} - E°_{anode} = +0.34$ V $-(-1.66$ V$) = +2.00$ V	1 point for correct $E°_{cell}$ (Must be consistent with part (a))

Question 6 (cont'd.)

(c) A cell is constructed based on the reaction in part (a) above. Label the metal used for the anode on the cell shown in the figure below.

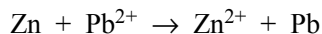| The metal is aluminum solid.	1 point for correct metal (Must be consistent with part (a))

(d) Of the compounds $NaOH$, CuS, and $NaNO_3$, which one is appropriate to use in a salt bridge? Briefly explain your answer, and for each of the other compounds, include a reason why it is not appropriate.

NaOH is not appropriate. The anion, OH^-, would migrate towards the anode. The OH^- would react with the Al^{3+} ion in solution. CuS is not appropriate. It is insoluble in water, so no ions would be available to migrate to the anode and cathode compartment to balance the charge. $NaNO_3$ is appropriate. It is soluble in water, and neither the cation nor the anion will react with the ions in the anode or cathode compartment.	1 point for correctly indicating whether each compound is appropriate, along with an explanation (3 points total)

Question 6 (cont'd.)

(e) Another standard cell is based on the following reaction.
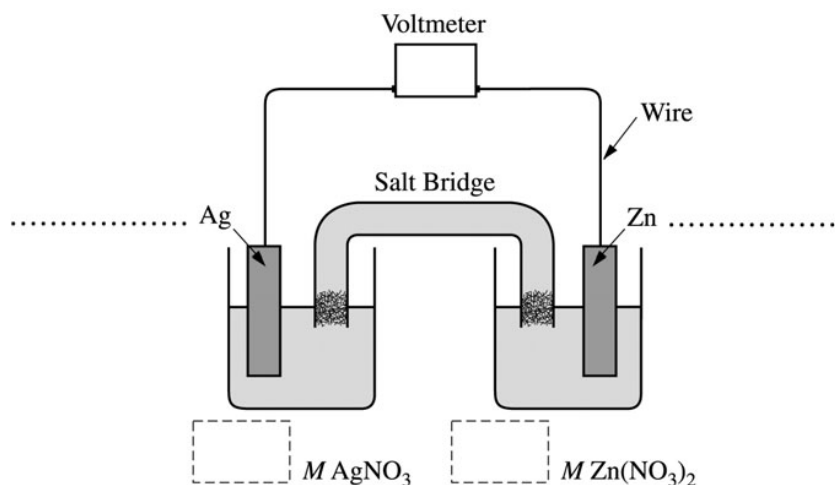
$$Zn + Pb^{2+} \rightarrow Zn^{2+} + Pb$$

If the concentration of Zn^{2+} is decreased from 1.0 M to 0.25 M, what effect does this have on the cell potential? Justify your answer.

$E_{cell} = E_{cell}^{\circ} - 0.059 \ln\left(\dfrac{[Zn^{2+}]}{[Pb^{2+}]}\right)$ If $[Zn^{2+}]$ is reduced, then the ratio $\left(\dfrac{[Zn^{2+}]}{[Pb^{2+}]}\right) < 1$, therefore $\ln\left(\dfrac{[Zn^{2+}]}{[Pb^{2+}]}\right) < 0$. Thus E_{cell} increases (becomes more positive).	1 point for correctly indicating how E_{cell} is affected 1 point for explanation in terms of Nernst equation and Q

Question 6

6. The following questions refer to the electrochemical cell shown in the diagram above.

(a) Write a balanced net ionic equation for the spontaneous reaction that takes place in the cell.

$Zn^{2+}(aq) + 2\,e^- \rightarrow Zn(s)$ $Ag^+(aq) + e^- \rightarrow Ag(s)$ $Zn(s) + 2\,Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2\,Ag(s)$	1 point for the balanced net-ionic equation 1 point for correct direction (reactants and products)

(b) Calculate the standard cell potential, $E°$, for the reaction in part (a).

$Zn^{2+}(aq) + 2\,e^- \rightarrow Zn(s) \qquad E° = -0.76$ V $Ag^+(aq) + e^- \rightarrow Ag(s) \qquad E° = +0.80$ V $E°_{cell} = E°_{cathode} - E°_{anode} = +0.80$ V $- (-0.76$ V$) = +1.56$ V	1 point for the correct, positive $E°_{cell}$

(c) In the diagram above,

(i) label the anode and the cathode on the dotted lines provided, and

The anode is the zinc metal electrode, the cathode is the silver metal electrode.	1 point for the correct metal for anode and cathode

Question 6 (cont'd.)

(ii) indicate, in the boxes below the half-cells, the concentration of $AgNO_3$ and the concentration of $Zn(NO_3)_2$ that are needed to generate $E°$

| $[AgNO_3] = [Zn(NO_3)_2] = 1\ M$ | 1 point for the correct concentration for the anodic chamber |
| | 1 point for the correct concentration for the cathodic chamber |

(d) How will the cell potential be affected if KI is added to the silver half-cell? Justify your answer.

| A precipitate will form as I^- ions react with Ag^+ ions in solution in cathode compartment. $[Ag^+]$ will be reduced, causing cell potential to decrease. | 1 point for correctly indicating a reaction occurs |
| | 1 point for indicating the cell potential decreases |

$$AgNO_3(s) \rightarrow Ag^+(aq) + NO_3^-(aq)$$
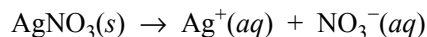
The dissolving of $AgNO_3(s)$ in pure water is represented by the equation above.

(a) Is ΔG for the dissolving of $AgNO_3(s)$ positive, negative, or zero? Justify your answer.

ΔG for the dissolving of $AgNO_3(s)$ is negative. Because $AgNO_3(s)$ is known to be soluble in water, the solution process must be spontaneous, therefore ΔG is negative.	One point is earned for the correct sign of ΔG and a correct explanation.

(b) Is ΔS for the dissolving of $AgNO_3(s)$ positive, negative, or zero? Justify your answer.

ΔS is positive because the solid reactant $AgNO_3(s)$ is more ordered than the aqueous ion products, $Ag^+(aq)$ and $NO_3^-(aq)$.	One point is earned for the correct sign of ΔS and a correct explanation.

(c) The solubility of $AgNO_3(s)$ increases with increasing temperature.

   (i) What is the sign of ΔH for the dissolving process? Justify your answer.

The sign of ΔH must be positive for the solubility of $AgNO_3$ to increase with increasing temperature. Solubility is an equilibrium process, and since increasing temperature (accomplished by adding heat) shifts the equilibrium towards the products side in the chemical equation, heat must be absorbed during the solution process. Therefore, the solution process is endothermic, and $\Delta H > 0$.	One point is earned for the correct sign of ΔH and a correct explanation.

   (ii) Is the answer you gave in part (a) consistent with your answers to parts (b) and (c)(i) ? Explain.

Yes. Although ΔH is positive, ΔS is also positive; thus ΔG can be negative because the value of the $T\Delta S$ term in the equation $\Delta G = \Delta H - T\Delta S$ is positive and can be greater than the value of the ΔH term. A positive number minus a greater positive number yields a negative number for the value of ΔG.	One point is earned for the correct sign and a correct explanation.

The compound NaI dissolves in pure water according to the equation $NaI(s) \rightarrow Na^+(aq) + I^-(aq)$. Some of the information in the table of standard reduction potentials given below may be useful in answering the questions that follow.

Half-reaction	$E°$ (V)
$O_2(g) + 4 H^+ + 4 e^- \rightarrow 2 H_2O(l)$	1.23
$I_2(s) + 2 e^- \rightarrow 2 I^-$	0.53
$2 H_2O(l) + 2 e^- \rightarrow H_2(g) + 2 OH^-$	–0.83
$Na^+ + e^- \rightarrow Na(s)$	–2.71

(d) An electric current is applied to a 1.0 M NaI solution.

   (i) Write the balanced oxidation half-reaction for the reaction that takes place.

$2 I^- \rightarrow I_2(s) + 2 e^-$	One point is earned for the correct half-reaction.

   (ii) Write the balanced reduction half-reaction for the reaction that takes place.

$2 H_2O(l) + 2 e^- \rightarrow H_2(g) + OH^-$	One point is earned for the correct half-reaction.

   (iii) Which reaction takes place at the anode, the oxidation reaction or the reduction reaction?

The oxidation half-reaction occurs at the anode.	One point is earned for the correct choice.

   (iv) All electrolysis reactions have the same sign for $\Delta G°$. Is the sign positive or negative? Justify your answer.

The sign of ΔG for all electrolysis reactions is positive. Because electrolysis reactions are non-spontaneous, energy in the form of applied electrical current (electrical work) must be applied to make the reaction occur.	One point is earned for the correct sign of ΔG and a correct explanation.

Electrolytic Cells

This page has no answers on it and serves as a divider between Galvanic and electrolytic types of electrochemical cells.

Explain each of the...

(a) When an aqueous solution of NaCl is electrolyzed, $Cl_2(g)$ is produced at the anode, but no $Na(s)$ is produced at the cathode.

(b) The mass of $Fe(s)$ produced when 1 faraday is used to reduce a solution of $FeSO_4$ is 1.5 times the mass of $Fe(s)$ produced when 1 faraday is used to reduce a solution of $FeCl_3$.

$$Zn + Pb^{2+}(1\text{-molar}) \rightarrow Zn^{2+}(1\text{-molar}) + Pb$$
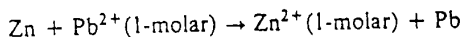
(c) The cell that utilizes the reaction above has a higher potential when $[Zn^{2+}]$ is decreased and $[Pb^{2+}]$ is held constant, but a lower potential when $[Pb^{2+}]$ is decreased and $[Zn^{2+}]$ is held constant.

(d) The cell that utilizes the reaction given in (c) has the same cell potential as another cell in which $[Zn^{2+}]$ and $[Pb^{2+}]$ are each 0.1-molar.

(a) Cl^- is more easily oxidized than H_2O (1 pt)

H_2O is more easily oxidized than Na^+ (1 pt)

- no 2nd pt is awarded for $H^+ \rightarrow H_2$ unless H_2O is implied

- no 2nd pt for $Na(s) + H_2O \rightarrow Na^+ + OH^-$ unless H_2 is indicated

(b) Fe^{+2} requires $2\mathcal{F}/mol\ Fe(s)$ or $1\mathcal{F} \rightarrow \frac{1}{2}\text{mol}\ Fe(s)$ (1 pt)

Fe^{+3} requires $3\mathcal{F}/mol\ Fe(s)$ $1\mathcal{F} \rightarrow \frac{1}{3}mol\ Fe(s)$

for equal #'s of $\mathcal{F}$ $\frac{1}{2} : \frac{1}{3} = 1.5$ to 1   (1 pt)

(or inverse relationship is clear)

- no 2nd pt unless flow of e^- to mass (mol) is clear and logically correct

(c) | LeChatelier's Argument | OR | Nernst Equation |

if $[Zn^{+2}] \downarrow$; rxn shifts $\rightarrow$ $E = E^\circ - RT\ln Q$ $Q = \frac{[Zn^{+2}]}{[Pb^{+2}]}$

(1 pt) ie, cell potential $\uparrow$ if $[Zn^{+2}] \downarrow$ $Q < 1$ $E > E^\circ$

(1 pt) if $[Pb^{+2}] \downarrow$; rxn shifts $\leftarrow$ if $[Pb^{+2}] \downarrow$, $Q > 1$ $E < E^\circ$

ie, cell potential $\downarrow$ - reasoning must indicate correct usage of equation

(d) $\frac{[Zn^{+2}]}{[Pb^{+2}]}$ does not change; regardless of values; ie,   OR $\frac{[Zn^{+2}]}{[Pb^{+2}]} = 1$ so $\ln Q = 0$

$E = E^\circ$ $E = E^\circ$

- no pt is awarded for just stating conc's are =

- ratio or proportion concept is required for 2nd pt

2. An unknown metal M forms a soluble compound, $M(NO_3)_2$.

  (a) A solution of $M(NO_3)_2$ is electrolyzed. When a constant current of 2.50 amperes is applied for 35.0 minutes, 3.06 grams of the metal M is deposited. Calculate the molar mass of M and identify the metal.

  (b) The metal identified in (a) is used with zinc to construct a galvanic cell, as shown below. Write the net ionic equation for the cell reaction and calculate the cell potential, $E°$.

Metal M   K⁺   Cl⁻   Zinc

1.0 M $M(NO_3)_2$ 1.0 M $ZnSO_4$
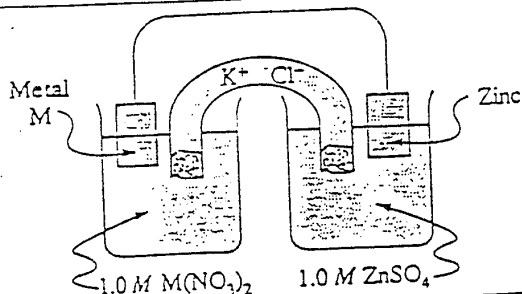
  (c) Calculate the value of the standard free energy change, $\Delta G°$, at 25°C for the reaction in (b).

  (d) Calculate the potential, E, for the cell shown in (b) if the initial concentration of $ZnSO_4$ is 0.10-molar, but the concentration of the $M(NO_3)_2$ solution remains unchanged.
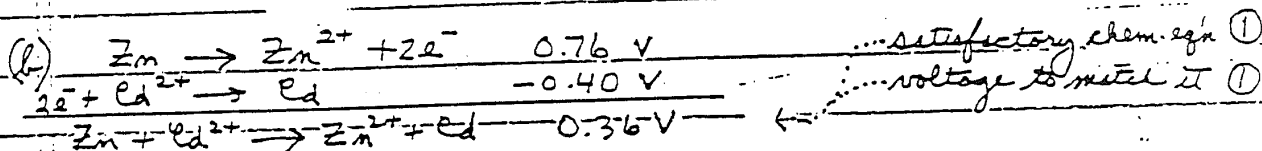
2. (a) 2.50 amp = 2.50 coul/sec

? coul = $2.50 \frac{coul}{sec} \times 35.0 \, min \times 60 \frac{sec}{min} = 5.25 \times 10^3$ coul   ①

? F = 5.25×10^3 coul $\times \frac{1 F}{96,500 \, coul} = 5.44 \times 10^{-2}$ F = 5.44×10^{-2} mol e⁻

? mol M = 5.44×10^{-2} mol e⁻ $\times \frac{1 \, mol \, M}{2 \, mol \, e⁻} = 2.72 \times 10^{-2}$ mol M   ①

? MM = $\frac{3.06 g}{2.72 \times 10^{-2} mol} = 112$ g/mol   ∴ metal is Cd   ①

(b) $Zn \rightarrow Zn^{2+} + 2e⁻$ 0.76 V   ...satisfactory chem. eq'n ①
$2e⁻ + Cd^{2+} \rightarrow Cd$ −0.40 V   ...voltage to match it ①
$Zn + Cd^{2+} \rightarrow Zn^{2+} + Cd$ 0.36 V

(c) $\Delta G° = -nFE° = -2(96,500 \frac{J}{V})(0.36 V) = -69,000 J = -69 \, kJ$

    ① for equation and ① substitution and correct answer   equation ①

(d) $E_{observed} = E° - \frac{0.059}{n} \log \frac{[Zn^{2+}]}{[Cd^{2+}]}$

    $= 0.36 V - \frac{0.059}{2} \log \frac{0.10}{1} = 0.36 V + 0.03 V = 0.39 V$ ①

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Question 3

(9 points)

(a) $2\,Cl^- \rightarrow Cl_2 + 2\,e^-$ (equation need not be balanced)   *1 point*
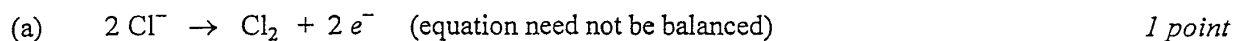
(b) $$\dfrac{0.250\ coul/sec \times 7{,}200\ sec}{96{,}500\ coul/mol\ e^-} = \dfrac{1{,}800\ coul}{96{,}500\ coul/mol\ e^-} = 0.01865\ mol\ e^-$$ *1 point*

$$mol\ Fe = \dfrac{0.521\ g\ Fe}{55.85\ g/mol\ Fe} = 0.00933\ mol\ Fe$$ *1 point*

$$\dfrac{mol\ e^-}{mol\ Fe} = \dfrac{1.865 \times 10^{-2}\ mol\ e^-}{9.33 \times 10^{-3}\ mol\ Fe} \approx 2\,e^-\ per\ Fe\ atom \Rightarrow FeCl_2$$ *1 point*

(c) $Fe^{2+} + 2\,Cl^- \rightarrow Fe + Cl_2$ *1 point*

<u>Notes:</u> "$FeCl_2(aq)$" accepted for reactants.
   Any balanced equation corresponding to answer in part (b) earns 1 point.

(d) moles Fe^{2+} = moles Cl_2 = $\mathbf{9.33 \times 10^{-3}\ mol\ Cl_2}$ *1 point*

$$V = \dfrac{nRT}{p} = \dfrac{0.00933\ mol\ Cl_2 \times 0.0821\ L \cdot atm \cdot mol^{-1} \cdot K^{-1} \times 298\ K}{\left(\frac{750}{760}\right)\ atm}$$

$$= \mathbf{0.231\ L\ (or\ 231\ mL)}$$ *1 point*

(e) $$\dfrac{3.00\ g\ Cl_2\ /\ 71\ g \cdot mol^{-1}}{3{,}600\ sec} = \dfrac{0.0423\ mol\ Cl_2}{3{,}600\ sec} = 1.17 \times 10^{-5}\ \dfrac{mol\ Cl_2}{sec}$$ *1 point*

$$current\ (in\ amperes) = \dfrac{2\ mol\ e^-}{mol\ Cl_2} \times 1.17 \times 10^{-5}\ \dfrac{mol\ Cl_2}{sec} \times \dfrac{96{,}500\ coul}{1\ mol\ e^-}$$

$$= \mathbf{2.27\ amp\ (coul/sec)}$$ *1 point*

<u>Alternate solution:</u> 0.00933 mole Cl_2/2 hrs = 0.662 g Cl_2/2 hrs = 0.331 g Cl_2/ hr

$$\dfrac{0.250\ amp}{0.331\ g\ Cl_2} = \dfrac{x}{3.00\ g\ Cl_2}$$

$$x = \dfrac{3.00\ g \times 0.250\ amp}{0.331\ g} = 2.27\ amp$$

Question 2
(10 points)

(a)   (i) $Zn(s) \rightarrow Zn^{2+}(aq) + 2\,e^-$ *1 pt.*

(ii) $Co^{2+}(aq) + Zn(s) \rightarrow Co(s) + Zn^{2+}(aq)$ *1 pt.*

(iii) $0.76\ V + (-0.28\ V) = 0.48\ V$ *1 pt.*

Note: phase designations not required in part (i) or part (ii)

(b)   (i) $\Delta G° = -nFE° = -2(96{,}500)(0.55V)$ *2 pts.*
                $= -1.1 \times 10^5\ J$ or $-1.1 \times 10^2\ kJ$

- First point earned for $n = 2$ (consistent use of $n = 4$ also accepted)
- Second point earned for negative sign, correct number (2 ± 1 sig. figs.), and appropriate units (kJ or J or kJ/mole or J/mole)

(ii) $\Delta G° = -RT\ \ln(K)$ *1 pt.*

$-1.1 \times 10^5\ J = -[8.31\ J\ mol^{-1}\ K^{-1}][298\ K][\ln(K)]$

$K = 2.0 \times 10^{19}$ (full credit also for correct use of $\log K = \dfrac{nE°}{0.0592}$)

(iii) $O_2 + 2\,H_2O \rightarrow 2\,H_2O_2$ $-0.55\ V$
        $O_2 + 4\,H^+ + 4\,e^- \rightarrow 2\,H_2O$ $1.23\ V$
        _____
        $2\,O_2 + 4\,H^+ + 4\,e^- \rightarrow 2\,H_2O_2$ $0.68\ V$ *2 pts.*
        $\Rightarrow O_2 + 2\,H^+ + 2\,e^- \rightarrow H_2O_2$ $0.68\ V$ (not required)
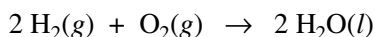
- Two points earned for correct voltage with supporting numbers (chemical equations not necessary)
- One point earned for correct chemical equations with incorrect voltage

(c) $3{,}600\ sec \times \dfrac{100\ C}{1\ sec} \times \dfrac{1\ mol\ e^-}{96{,}500\ C} \times \dfrac{1\ mol\ Cu}{2\ mol\ e^-} \times \dfrac{63.55\ g\ Cu}{1\ mol\ Cu} = 119\ g\ Cu$ *2 pts.*

- Two points earned for correct answer (3 ± 1 sig. figs.)
- One point earned for any two of these steps: (amp)(sec) $\Rightarrow$ coulombs
                                                    coulombs $\Rightarrow$ mol e^-
                                                    mol e^- $\Rightarrow$ mol Cu
                                                    mol Cu $\Rightarrow$ g Cu

Question 3

$$2 \, H_2(g) + O_2(g) \rightarrow 2 \, H_2O(l)$$

In a hydrogen-oxygen fuel cell, energy is produced by the overall reaction represented above.

(a) When the fuel cell operates at 25°C and 1.00 atm for 78.0 minutes, 0.0746 mol of $O_2(g)$ is consumed. Calculate the volume of $H_2(g)$ consumed during the same time period. Express your answer in liters measured at 25°C and 1.00 atm.

$(0.0746 \text{ mol } O_2) \times \dfrac{2 \text{ mol } H_2}{1 \text{ mol } O_2} = 0.149 \text{ mol } H_2$ $V = \dfrac{n_{H_2} RT}{P} = \dfrac{(0.149 \text{ mol } H_2)(0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1})(298 \text{ K})}{1.00 \text{ atm}}$ $= 3.65 \text{ L } H_2$	One point is earned for calculation of moles of H_2. One point is earned for substitution into $PV = nRT$. One point is earned for the answer.

(b) Given that the fuel cell reaction takes place in an acidic medium,

   (i) write the two half reactions that occur as the cell operates,

$O_2 + 4 \, H^+ + 4 \, e^- \rightarrow 2 \, H_2O$ $H_2 \rightarrow 2 \, H^+ + 2 \, e^-$	One point is earned for each of the two half reactions.

   (ii) identify the half reaction that takes place at the cathode, and

$O_2 + 4 \, H^+ + 4 \, e^- \rightarrow 2 \, H_2O$	One point is earned for either the equation of the correct half reaction, or for indicating "the reduction half reaction" if the correct equation is given in (b)(i).

   (iii) determine the value of the standard potential, $E°$, of the cell.

$E° = 1.23 \text{V} + 0.00 \text{ V} = 1.23 \text{ V}$	One point is earned for the standard potential.

Question 3 (continued)

(c) Calculate the charge, in coulombs, that passes through the cell during the 78.0 minutes of operation as described in part (a).

$$(0.0746 \text{ mol } O_2) \times \frac{4 \text{ mol } e^-}{1 \text{ mol } O_2} \times \frac{96,500 \text{ C}}{1 \text{ mol } e^-} = 2.88 \times 10^4 \text{ C}$$	One point is earned for the stoichiometry. One point is earned for the answer.