Question 2

2. Answer the following questions about voltaic cells.

(a) A voltaic cell is set up using Al/Al^{3+} as one half-cell and Sn/Sn^{2+} as the other half-cell. The half-cells contain equal volumes of solutions and are at standard conditions.

(i) Write the balanced net-ionic equation for the spontaneous cell reaction.

$3\,Sn^{2+} + 2\,Al \rightarrow 3\,Sn + 2\,Al^{3+}$	One point is earned for the correct direction. One point is earned for the balanced net-ionic equation.

(ii) Determine the value, in volts, of the standard potential, $E°$, for the spontaneous cell reaction.

$E° = -0.14\text{ V} - (-1.66\text{ V}) = 1.52\text{ V}$ (or, 1.52 J C^{-1})	One point is earned for the correct answer. (Potential <u>must</u> be positive.)

(iii) Calculate the value of the standard free-energy change, $\Delta G°$, for the spontaneous cell reaction. Include units with your answer.

$\Delta G° = -nFE° = -\frac{6\text{ mol } e^-}{1\text{ mol}} \times \frac{96{,}500\text{ C}}{1\text{ mol } e^-} \times (1.52\text{ J C}^{-1})$ $= -8.80 \times 10^5\text{ J mol}^{-1}$ (or -880 kJ mol^{-1})	One point is earned for indicating the correct mol e^- to mol reaction ratio. One point is earned for the correct answer with correct units.

(iv) If the cell operates until $[Al^{3+}]$ is 1.08 M in the Al/Al^{3+} half-cell, what is $[Sn^{2+}]$ in the Sn/Sn^{2+} half-cell?

change in $[Sn^{2+}] = \frac{0.08\text{ mol } Al^{3+}}{1\text{ L}} \times \frac{3\text{ mol } Sn^{2+}}{2\text{ mol } Al^{3+}} = \frac{0.12\text{ mol } Sn^{2+}}{1\text{ L}}$ $[Sn^{2+}] = 1.00\text{ mol L}^{-1} - 0.12\text{ mol L}^{-1} = 0.88\text{ mol L}^{-1}$	One point is earned for the correct answer.

(b) In another voltaic cell with Al/Al^{3+} and Sn/Sn^{2+} half-cells, $[Sn^{2+}]$ is 0.010 M and $[Al^{3+}]$ is 1.00 M. Calculate the value, in volts, of the cell potential, E_{cell}, at 25°C.

$E_{cell} = 1.52\text{ V} - \frac{0.0592}{6}\log\frac{(1.00)^2}{(0.010)^3}$ $= 1.52\text{ V} - 0.0592\text{ V} = 1.46\text{ V}$	Answers must be consistent with part (a)(i). One point is earned for the proper exponents. One point is earned for the correct substitution of concentrations. One point is earned for the correct answer.

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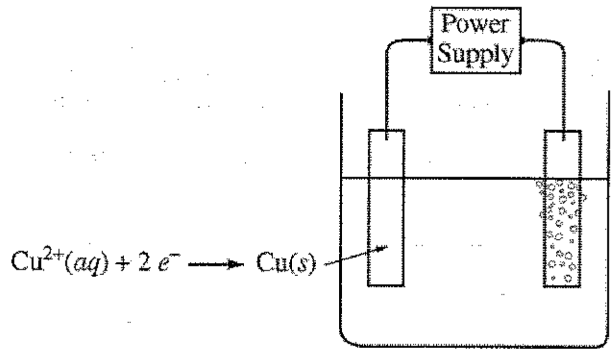
Question 3



An external direct-current power supply is connected to two platinum electrodes immersed in a beaker containing 1.0 *M* $CuSO_4(aq)$ at 25°C, as shown in the diagram above. As the cell operates, copper metal is deposited onto one electrode and $O_2(g)$ is produced at the other electrode. The two reduction half-reactions for the overall reaction that occurs in the cell are shown in the table below.

Half-Reaction	$E°$(V)
$O_2(g) + 4\,H^+(aq) + 4\,e^- \rightarrow 2\,H_2O(l)$	+1.23
$Cu^{2+}(aq) + 2\,e^- \rightarrow Cu(s)$	+0.34

(a) On the diagram, indicate the direction of electron flow in the wire.

The electron flow in the wire is from the right toward the left (counterclockwise).	One point is earned for the correct direction.

(b) Write a balanced net ionic equation for the electrolysis reaction that occurs in the cell.

$2\,H_2O(l) + 2\,Cu^{2+}(aq) \rightarrow 4\,H^+(aq) + 2\,Cu(s) + O_2(g)$	One point is earned for all three products. One point is earned for balancing the equation.

(c) Predict the algebraic sign of $\Delta G°$ for the reaction. Justify your prediction.

The sign of $\Delta G°$ would be positive because the reaction is NOT spontaneous.	One point is earned for indicating that $\Delta G°$ is greater than zero <u>and</u> supplying a correct explanation.

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Question 3 (continued)

(d) Calculate the value of $\Delta G°$ for the reaction.

$E° = -1.23\text{ V} + 0.34\text{ V} = -0.89\text{ V} = -0.89\text{ J C}^{-1}$ $\Delta G° = -n\mathcal{F}E° = -4(96{,}500\text{ C mol}^{-1})(-0.89\text{ J C}^{-1})$ $= +340{,}000\text{ J mol}^{-1} = +340\text{ kJ mol}^{-1}$	One point is earned for calculating $E°$. One point is earned for calculating $\Delta G°$ (consistent with the calculated $E°$).

An electric current of 1.50 amps passes through the cell for 40.0 minutes.

(e) Calculate the mass, in grams, of the Cu(s) that is deposited on the electrode.

$q = (1.50\text{ C s}^{-1})(40.0\text{ min}) \times \frac{60\text{ s}}{1\text{ minute}} = 3{,}600\text{ C}$ $\text{mass Cu} = (3{,}600\text{ C}) \times \frac{1\text{ mol } e^-}{96{,}500\text{ C}} \times \frac{1\text{ mol Cu}}{2\text{ mol } e^-} \times \frac{63.55\text{ g Cu}}{1\text{ mol Cu}}$ $= 1.19\text{ g Cu}$	One point is earned for calculating q. One point is earned for calculating the mass of copper deposited. ***OR*** Two points are earned for calculating the mass of copper in one step.

(f) Calculate the dry volume, in liters measured at 25°C and 1.16 atm, of the $O_2(g)$ that is produced.

$n_{O_2} = (1.19\text{ g Cu}) \times \frac{1\text{ mol Cu}}{63.55\text{ g Cu}} \times \frac{1\text{ mol O}_2}{2\text{ mol Cu}} = 0.00936\text{ mol O}_2$ $V = \frac{nRT}{P} = \frac{(0.00936\text{ mol})(0.0821\text{ L atm mol}^{-1}\text{ K}^{-1})(298\text{ K})}{1.16\text{ atm}}$ $= 0.197\text{ L}$	One point is earned for calculating the number of moles of O_2. One point is earned for calculating V (consistent with previous calculations).

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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.

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

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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.

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Question 3

Answer the following questions related to chemical reactions involving nitrogen monoxide, NO(g).

The reaction between solid copper and nitric acid to form copper(II) ion, nitrogen monoxide gas, and water is represented by the following equation.

$$3\,Cu(s) + 2\,NO_3^-(aq) + 8\,H^+(aq) \rightarrow 3\,Cu^{2+}(aq) + 2\,NO(g) + 4\,H_2O(l) \qquad E° = +0.62\text{ V}$$

(a) Using the information above and in the table below, calculate the standard reduction potential, $E°$, for the reduction of NO_3^- in acidic solution.

Half-Reaction	Standard Reduction Potential, $E°$
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	+0.34 V
$NO_3^-(aq) + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(l)$	?

$E°_{rxn} = E°_{NO_3^-} - E°_{Cu^{2+}} = E°_{NO_3^-} - 0.34\text{ V} = 0.62\text{ V}$ $\Rightarrow E°_{NO_3^-} = 0.62\text{ V} + 0.34\text{ V} = 0.96\text{ V}$	One point is earned for the correct calculation of the standard reduction potential.

(b) Calculate the value of the standard free energy change, $\Delta G°$, for the overall reaction between solid copper and nitric acid.

$\Delta G° = -n\mathcal{F}E° = -(6)(96{,}500\text{ C mol}^{-1})(0.62\text{ V})$ $= -360{,}000\text{ J mol}^{-1} = -360\text{ kJ mol}^{-1}$	One point is earned for the correct value of n, the number of moles of electrons. One point is earned for calculating the correct value of $\Delta G°$, with correct sign and consistent units.

(c) Predict whether the value of the standard entropy change, $\Delta S°$, for the overall reaction is greater than 0, less than 0, or equal to 0. Justify your prediction.

$\Delta S° > 0$. Even though there is a loss of 7 moles of ions in solution, the value of $\Delta S°$ for the overall reaction will be greater than zero because two moles of NO gas will be produced (there are no gaseous reactants).	One point is earned for the correct answer with a justification that is based on the gaseous state of one of the products.

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Question 3 (continued)

Nitrogen monoxide gas, a product of the reaction above, can react with oxygen to produce nitrogen dioxide gas, as represented below.

$$2\,NO(g) + O_2(g) \rightarrow 2\,NO_2(g)$$

A rate study of the reaction yielded the data recorded in the table below.

Experiment	Initial Concentration of NO (mol L^{-1})	Initial Concentration of O_2 (mol L^{-1})	Initial Rate of Formation of NO_2 (mol L^{-1} s^{-1})
1	0.0200	0.0300	8.52×10^{-2}
2	0.0200	0.0900	2.56×10^{-1}
3	0.0600	0.0300	7.67×10^{-1}

(d) Determine the order of the reaction with respect to each of the following reactants. Give details of your reasoning, clearly explaining or showing how you arrived at your answers.

(i) NO

Comparing experiments 1 and 3, the tripling of the initial concentration of NO while the initial concentration of oxygen remained constant at 0.0300 mol L^{-1} resulted in a nine-fold increase in the initial rate of formation of NO_2. Since $9 = 3^2$, the reaction is second order with respect to NO.	One point is earned for the correct answer with justification.

(ii) O_2

Comparing experiments 1 and 2, the tripling of the initial concentration of O_2 while the initial concentration of NO remained constant at 0.0200 mol L^{-1} resulted in a tripling in the initial rate of formation of NO_2. Since $3 = 3^1$, the reaction is first order with respect to O_2.	One point is earned for the correct answer with justification.

(e) Write the expression for the rate law for the reaction as determined from the experimental data.

rate = $k\,[NO]^2[O_2]$	One point is earned for the correct expression for the rate law.

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Question 3 (continued)

(f) Determine the value of the rate constant for the reaction, clearly indicating the units.

Because the coefficient for NO_2 in the balanced equation is 2, the rate of the reaction is defined as $\frac{1}{2}$ the rate of the appearance of NO_2. From part (e) above, $k = \dfrac{\text{reaction rate}}{[NO]^2[O_2]}$ $= \dfrac{\left(\frac{1}{2}\right)(\text{rate of formation of } NO_2)}{[NO]^2[O_2]}$ Substituting data from experiment 1, $k = \dfrac{\left(\frac{1}{2}\right)(8.52 \times 10^{-2}\ \text{mol L}^{-1}\,\text{s}^{-1})}{(0.0200\ \text{mol L}^{-1})^2(0.0300\ \text{mol L}^{-1})}$ $= 3.55 \times 10^3\ \text{L}^2\,\text{mol}^{-2}\,\text{s}^{-1}$	One point is earned for calculating the correct value of the rate constant. One point is earned for including the correct units.

Note: a rate constant value of $7.10 \times 10^3\ \text{L}^2\,\text{mol}^{-2}\,\text{s}^{-1}$ earns the point if the rate of reaction is assumed to be the same as the rate of formation of NO_2.

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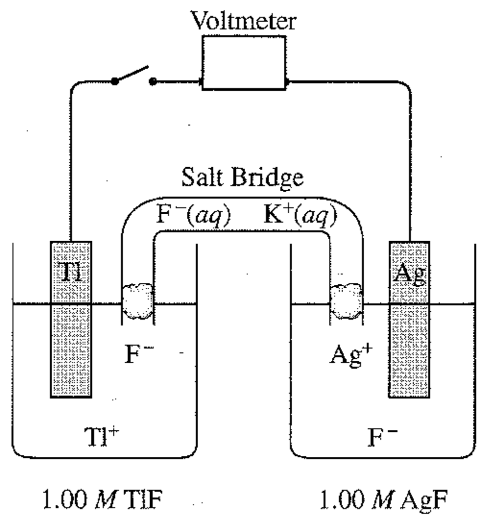
Question 6 (9 points)

Answer the following questions about electrochemical cells.



It is observed that when silver metal is placed in aqueous thallium(I) fluoride, TlF, no reaction occurs. When the switch is closed in the cell represented above, the voltage reading is +1.14 V.

(a) Write the reduction half-reaction that occurs in the cell.

$Ag^+ + e^- \rightarrow Ag$	One point is earned for the correct equation.

(b) Write the equation for the overall reaction that occurs in the cell.

$Tl + Ag^+ \rightarrow Tl^+ + Ag$	One point is earned for the correct equation.

(c) Identify the anode in the cell. Justify your answer.

The anode is where oxidation occurs. In the overall reaction Tl is oxidized to Tl^+, so the anode is the Tl electrode in the left cell.	One point is earned for the correct answer with justification.

(d) On the diagram above, use an arrow to clearly indicate the direction of electron flow as the cell operates.

The arrow should show electron flow in the direction from the Tl electrode through the wire to the Ag electrode.	One point is earned for a correct arrow.

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Question 6 (continued)

(e) Calculate the value of the standard reduction potential for the Tl^{+}/Tl half-reaction.

$E^{\circ}_{cell} = E^{\circ}_{red} - E^{\circ}_{ox}$ $+1.14\text{ V} = +0.80\text{V} - E^{\circ}_{ox}$ $E^{\circ}_{ox} = \mathbf{-0.34\text{ V}}$	One point is earned for the correct setup. One point is earned for the correct answer.

The standard reduction potential, E°, of the reaction $Pt^{2+} + 2\,e^{-} \rightarrow Pt$ is 1.20 V.

(f) Assume that electrodes of pure Pt, Ag, and Ni are available as well as 1.00 *M* solutions of their salts. Three different electrochemical cells can be constructed using these materials. Identify the two metals that when used to make an electrochemical cell would produce the cell with the largest voltage. Explain how you arrived at your answer.

E°(V) $Ni^{2+} + 2\,e^{-} \rightarrow Ni$ −0.25 $Ag^{+} + e^{-} \rightarrow Ag$ 0.80 $Pt^{2+} + 2\,e^{-} \rightarrow Pt$ 1.20 $E^{\circ}_{cell} = E^{\circ}_{red} - E^{\circ}_{ox}$ The two metals that yield the largest E°_{cell} are those with the biggest difference in E°, namely, Pt and Ni (see E°_{cell} calculation below). $E^{\circ}_{cell} = +1.20 - (-0.25) = +1.45\text{ V}$	One point is earned for the correct answer with justification.

(g) Predict whether Pt metal will react when it is placed in 1.00 *M* $AgNO_3(aq)$. Justify your answer.

When Pt metal is added to 1.00 *M* $AgNO_3$, the only redox reaction that could occur would be for Pt to become oxidized as Ag^{+} is reduced. $E^{\circ}_{cell} = E^{\circ}_{red} - E^{\circ}_{ox} = +0.80\text{ V} - (+1.20\text{ V}) = -0.40\text{ V}$ Because E°_{cell} for that reaction is negative, no reaction will occur.	One point is earned for comparing E° values. One point is earned for the correct interpretation.

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Question 6
(9 points)

$$2\ Al(s) + 3\ Zn^{2+}(aq) \rightarrow 2\ Al^{3+}(aq) + 3\ Zn(s)$$

Respond to the following statements and questions that relate to the species and the reaction represented above.

(a) Write the complete electron configuration (e.g., $1s^2 2s^2 \ldots$) for Zn^{2+}.

$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$	One point is earned for the correct configuration.

(b) Which species, Zn or Zn^{2+}, has the greater ionization energy? Justify your answer.

Zn^{2+} has the greater ionization energy. The electron being removed from Zn^{2+} experiences a larger effective nuclear charge than the electron being removed from Zn because Zn^{2+} has two fewer electrons shielding the nucleus. ***OR*** It takes more energy to remove a negatively charged electron from a positive ion than from a neutral atom.	One point is earned for identifying Zn^{2+} with justification.

(c) Identify the species that is oxidized in the reaction.

Al(*s*)	One point is earned for identifying Al.

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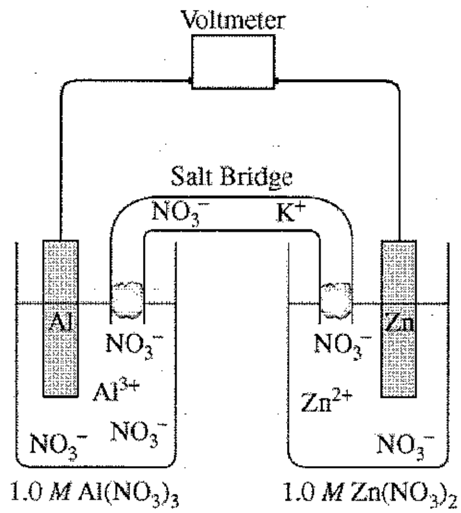
Question 6 (continued)

The diagram below shows a galvanic cell based on the reaction. Assume that the temperature is 25°C.



(d) The diagram includes a salt bridge that is filled with a saturated solution of KNO_3. Describe what happens in the salt bridge as the cell operates.

As the cell operates, NO_3^- ions flow toward the Al half-cell and K^+ ions flow toward the Zn half-cell.	One point is earned for correctly indicating the direction of ion flow.

(e) Determine the value of the standard voltage, $E°$, for the cell.

$E° = (-0.76 \text{ V}) - (-1.66 \text{ V}) = 0.90 \text{ V}$	One point is earned for the correct $E°$.

(f) Indicate whether the value of the standard free-energy change, $\Delta G°$, for the cell reaction is positive, negative, or zero. Justify your answer.

$\Delta G°$ is negative since $E°$ is positive and $\Delta G° = -n\mathcal{F}E°$. ***OR*** $\Delta G°$ must be negative because the reaction is spontaneous under standard conditions.	One point is earned for indicating that $\Delta G°$ is negative. One point is earned for a correct justification.

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Question 6 (continued)

(g) If the concentration of $Al(NO_3)_3$ in the $Al(s)/Al^{3+}(aq)$ half-cell is lowered from 1.0 M to 0.01 M at 25°C, does the cell voltage increase, decrease, or remain the same? Justify your answer.

Lowering $[Al^{3+}]$ causes an increase in the cell voltage. The value of Q will fall below 1.0 and the log term in the Nernst equation will become negative. This causes the value of E_{cell} to become more positive. ***OR*** A decrease in a product concentration will increase the spontaneity of the reaction, increasing the value of E_{cell}.	One point is earned for indicating that E_{cell} increases. One point is earned for the correct justification.

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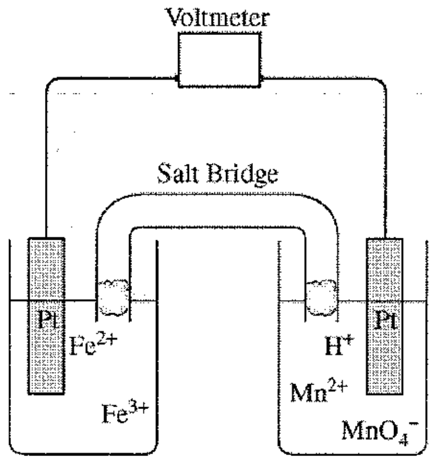
Question 2
(10 points)



$$5\ Fe^{2+}(aq) + MnO_4^-(aq) + 8\ H^+(aq) \rightarrow 5\ Fe^{3+}(aq) + Mn^{2+}(aq) + 4\ H_2O(l)$$

A galvanic cell and the balanced equation for the spontaneous cell reaction are shown above. The two reduction half-reactions for the overall reaction that occurs in the cell are shown in the table below.

Half-Reaction	$E°$ (V) at 298 K
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+0.77
$MnO_4^-(aq) + 8\ H^+(aq) + 5\ e^- \rightarrow Mn^{2+}(aq) + 4\ H_2O(l)$	+1.49

(a) On the diagram, clearly label the cathode.

The electrode in the beaker on the right should be labeled.	One point is earned for correct identification of the cathode.

(b) Calculate the value of the standard potential, $E°$, for the spontaneous cell reaction.

$E_{cell} = 1.49 - 0.77 = 0.72$ V	One point is earned for the correct numerical answer.

(c) How many moles of electrons are transferred when 1.0 mol of $MnO_4^-(aq)$ is consumed in the overall cell reaction?

5.0 moles of electrons are transferred.	One point is earned for the correct numerical answer.

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Question 2 (continued)

(d) Calculate the value of the equilibrium constant, K_{eq}, for the cell reaction at 25°C. Explain what the magnitude of K_{eq} tells you about the extent of the reaction.

$\log K_{eq} = \frac{nE}{0.0592} = \frac{5 \times 0.72}{0.0592} = 61$ $K_{eq} = 6.5 \times 10^{60}$ Because the magnitude of K_{eq} is very large, the extent of the cell reaction is also very large and the reaction goes essentially to completion.	One point is earned for the correct substitution. One point is earned for the correct numerical answer. One point is earned for an explanation.

Three solutions, one containing $Fe^{2+}(aq)$, one containing $MnO_4^-(aq)$, and one containing $H^+(aq)$, are mixed in a beaker and allowed to react. The initial concentrations of the species in the mixture are 0.60 M $Fe^{2+}(aq)$, 0.10 M $MnO_4^-(aq)$, and 1.0 M $H^+(aq)$.

(e) When the reaction mixture has come to equilibrium, which species has the higher concentration, $Mn^{2+}(aq)$ or $MnO_4^-(aq)$? Explain.

$[Mn^{2+}(aq)]$ will be greater than $[MnO_4^-(aq)]$ because: (1) as indicated in part (d), the reaction essentially goes to completion, and (2) there is more than sufficient Fe^{2+} and H^+ to react completely with the MnO_4^-. $[MnO_4^-(aq)]$ at equilibrium is essentially zero.	One point is earned for the choice of Mn^{2+} with the explanation including only item (1). One point is earned for including item (2) in the explanation.

(f) When the reaction mixture has come to equilibrium, what are the molar concentrations of $Fe^{2+}(aq)$ and $Fe^{3+}(aq)$?

At equilibrium, $[Fe^{2+}(aq)] = [Fe^{2+}(aq)]_{initial} - 5[MnO_4^-(aq)]_{reacted}$ $= 0.60 - 5(0.10) = 0.10\ M$ $[Fe^{3+}(aq)] = 5 \times [MnO_4^-(aq)]_{reacted}$ $= 5\,(0.10) = 0.50\ M$	One point is earned for a correct setup (including a correct setup for an equilibrium calculation). One point is earned for correct numerical answers.

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Question 3

Hydrogen gas burns in air according to the equation below.

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

(a) Calculate the standard enthalpy change, ΔH°_{298}, for the reaction represented by the equation above. (The molar enthalpy of formation, ΔH°_f, for $H_2O(l)$ is -285.8 kJ mol^{-1} at 298 K.)

$\Delta H^\circ_{298} = [2\,(-285.8)] - [2(0) + 1(0)] = -571.6$ kJ mol^{-1}	1 point is earned for the correct answer.

(b) Calculate the amount of heat, in kJ, that is released when 10.0 g of $H_2(g)$ is burned in air.

$q = 10\text{ g H}_2 \times \dfrac{1\text{ mol H}_2}{2.016\text{ g H}_2} \times \dfrac{285.8\text{ kJ}}{1\text{ mol H}_2} = 1.42 \times 10^3\text{ kJ}$	1 point is earned for the correct setup. 1 point is earned for the correct answer.

(c) Given that the molar enthalpy of vaporization, ΔH°_{vap}, for $H_2O(l)$ is 44.0 kJ mol^{-1} at 298 K, what is the standard enthalpy change, ΔH°_{298}, for the reaction $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$?

$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$ $\quad -571.6$ kJ $2\ H_2O(l) \rightarrow 2\ H_2O(g)$ $\quad +2(44.0)$ kJ ——————————————— $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$ $\quad -483.6$ kJ	1 point is earned for the correct answer.

A fuel cell is an electrochemical cell that converts the chemical energy stored in a fuel into electrical energy. A cell that uses H_2 as the fuel can be constructed based on the following half-reactions.

Half-reaction	E° (298 K)
$2\ H_2O(l) + O_2(g) + 4\ e^- \rightarrow 4\ OH^-(aq)$	0.40 V
$2\ H_2O(l) + 2\ e^- \rightarrow H_2(g) + 2\ OH^-(aq)$	-0.83 V

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Question 3 (continued)

(d) Write the equation for the overall cell reaction.

$2\,H_2O(l) + O_2(g) + 4\,e^- \rightarrow 4\,OH^-(aq)$ $2\,H_2(g) + 4\,OH^-(aq) \rightarrow 4\,H_2O(l) + 4\,e^-$ ——————————————— $2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(l)$	1 point is earned for the correct equation.

(e) Calculate the standard potential for the cell at 298 K.

$E° = 0.40\text{ V} - (-0.83\text{ V}) = 1.23\text{ V}$	1 point is earned for the correct answer.

(f) Assume that 0.93 mol of $H_2(g)$ is consumed as the cell operates for 600. seconds.

(i) Calculate the number of moles of electrons that pass through the cell.

$0.93\text{ mol } H_2 \times \dfrac{2\text{ mol } e^-}{1\text{ mol } H_2} = 1.9\text{ mol } e^-$	1 point is earned for the correct answer.

(ii) Calculate the average current, in amperes, that passes through the cell.

$1.9\text{ mol } e^- \times \dfrac{96{,}500\text{ C}}{1\text{ mol } e^-} = 1.8 \times 10^5\text{ C}$	1 point is earned for calculation of the charge in coulombs.
$I = \dfrac{q}{t} = \dfrac{1.8 \times 10^5\text{ C}}{600.\text{ s}} = 3.0 \times 10^2\text{ amps}$	1 point is earned for calculation of the current in amperes.

(g) Some fuel cells use butane gas, C_4H_{10}, rather than hydrogen gas. The overall reaction that occurs in a butane fuel cell is $2\,C_4H_{10}(g) + 13\,O_2(g) \rightarrow 8\,CO_2(g) + 10\,H_2O(l)$. What is one environmental advantage of using fuel cells that are based on hydrogen rather than on hydrocarbons such as butane?

Hydrogen fuel cells produce only water as a product, unlike fuel cells that use hydrocarbons, which release carbon dioxide. Carbon dioxide contributes to global warming via the enhanced atmospheric greenhouse effect.	1 point is earned for an acceptable environmental advantage.

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Question 6
(9 points)

In a laboratory experiment, Pb and an unknown metal Q were immersed in solutions containing aqueous ions of unknown metals Q and X. The following reactions summarize the observations.

Observation 1: $Pb(s) + X^{2+}(aq) \rightarrow Pb^{2+}(aq) + X(s)$
Observation 2: $Q(s) + X^{2+}(aq) \rightarrow$ no reaction
Observation 3: $Pb(s) + Q^{2+}(aq) \rightarrow Pb^{2+}(aq) + Q(s)$

(a) On the basis of the reactions indicated above, arrange the three metals, Pb, Q, and X, in order from least reactive to most reactive on the lines provided below.

____Q____ , ____X____ , ____Pb____

least reactive metal ... most reactive metal

Q, X, Pb	2 points are earned for the correctly ordered relationship. (1 point earned for Q, Pb, X or X, Q, Pb)

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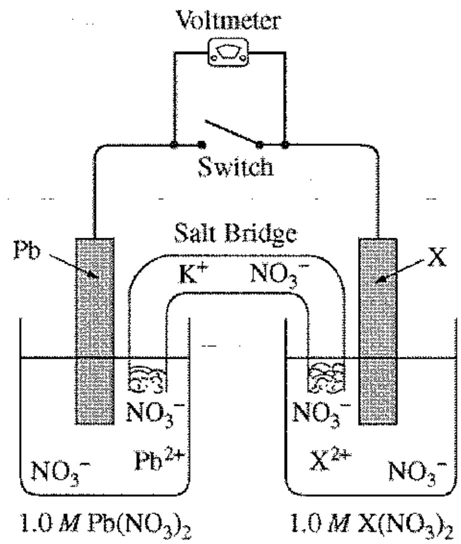
Question 6 (continued)

The diagram below shows an electrochemical cell that is constructed with a Pb electrode immersed in 100. mL of 1.0 *M* $Pb(NO_3)_2(aq)$ and an electrode made of metal X immersed in 100. mL of 1.0 *M* $X(NO_3)_2(aq)$. A salt bridge containing saturated aqueous KNO_3 connects the anode compartment to the cathode compartment. The electrodes are connected to an external circuit containing a switch, which is open. When a voltmeter is connected to the circuit as shown, the reading on the voltmeter is 0.47 V. When the switch is closed, electrons flow through the switch from the Pb electrode toward the X electrode.



(b) Write the equation for the half-reaction that occurs at the anode.

$Pb(s) \rightarrow Pb^{2+}(aq) + 2\,e^-$	1 point is earned for the correct equation.

(c) The value of the standard potential for the cell, $E°$, is 0.47 V.

(i) Determine the standard reduction potential for the half-reaction that occurs at the cathode.

$E°_{cell} = E°_{cathode} - E°_{anode}$ $E°_{cathode} = E°_{cell} + E°_{anode}$ $E°_{cathode} = 0.47 + (-0.13) = 0.34$ V	1 point is earned for the calculated reduction potential with mathematical justification.

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Question 6 (continued)

(ii) Determine the identity of metal X.

The metal is copper.	1 point is earned for identification of the metal.

(d) Describe what happens to the mass of each electrode as the cell operates.

The mass of the Pb electrode decreases and the mass of the Cu electrode increases.	1 point is earned for both descriptions.

(e) During a laboratory session, students set up the electrochemical cell shown above. For each of the following three scenarios, choose the correct value of the cell voltage and justify your choice.

(i) A student bumps the cell setup, resulting in the salt bridge losing contact with the solution in the cathode compartment. Is V equal to 0.47 or is V equal to 0? Justify your choice.

V = 0 V. The transfer of ions through the salt bridge will stop. A charge imbalance between the half-cells will prevent electrons from flowing through the wire.	1 point is earned for the correct choice with an appropriate explanation.

(ii) A student spills a small amount of 0.5 *M* $Na_2SO_4(aq)$ into the compartment with the Pb electrode, resulting in the formation of a precipitate. Is V less than 0.47 or is V greater than 0.47 ? Justify your choice.

V > 0.47 V. The sulfate ion will react with the Pb^{2+} ion to form a precipitate. This results in a thermodynamically favored anode half-cell reaction and hence a larger potential difference. The choice may also be justified using the Nernst equation. $E_{cell} = E^\circ_{cell} - \left(\frac{RT}{nF}\right)\ln\frac{[Pb^{2+}]}{[Cu^{2+}]}$ Decreasing the $[Pb^{2+}]$ will increase the cell voltage.	1 point is earned for the correct choice with an appropriate explanation.

(iii) After the laboratory session is over, a student leaves the switch closed. The next day, the student opens the switch and reads the voltmeter. Is V less than 0.47 or is V equal to 0.47? Justify your choice.

V < 0.47 V. Over time, $[Pb^{2+}]$ increases and $[Cu^{2+}]$ decreases, making both half-cell reactions less thermodynamically favorable. The choice may also be justified using the Nernst equation. Increasing $[Pb^{2+}]$ and decreasing $[Cu^{2+}]$ decreases the cell voltage. The choice may also be justified by stating that the voltage is zero as a result of the establishment of equilibrium.	1 point is earned for the correct choice with an appropriate explanation.

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