2002 (B)

Question 7 page 2

(a)

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

Red → $$Ni^{2+} + 2e^- \rightarrow Ni(s) \quad -.25$$

Ox → $$Zn(s) \rightarrow Zn^{2+} + 2e^- \quad +.76$$

$$E^\circ_{cell} = .51\ V$$

$M = Zn$

$M^{2+} = Zn^{2+}$ (1 M)

(b) Ni is the cathode. Reduction takes place at the cathode.

$$Ni^{2+} + 2e^- \rightarrow Ni(s)$$

(c)

$$E_{cell} = E^\circ_{cell} - \frac{RT}{nF} \ln Q$$

$$= E^\circ_{cell} - \frac{RT}{nF} \ln \frac{.1}{1}$$

$Q < 1$

$\ln Q = (-)$

$$= (+) - (+)(-)$$

$$= (+) - (-) \rightarrow E_{cell} > E^\circ_{cell}$$

Cell voltage would increase.

(d) No voltage if salt bridge is removed.

Question 2
page 3

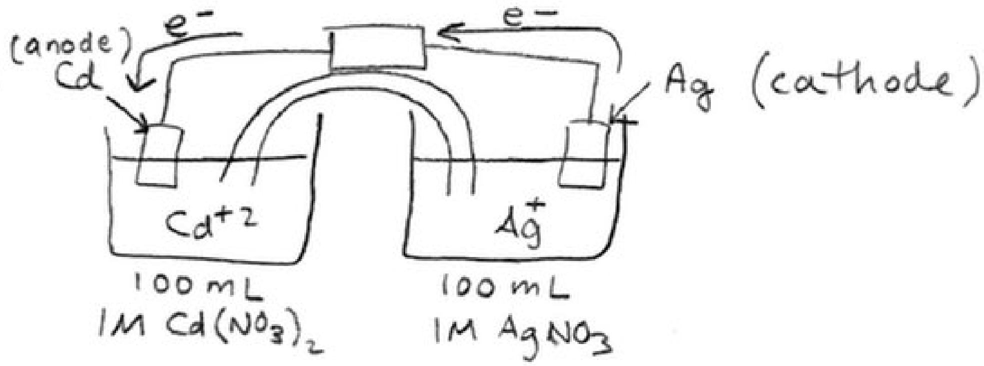
2 (a) red $Cd^{+2} + 2e^- \rightarrow Cd$ $\mathcal{E}^\circ = .80V$

$Ag^+ + e^- \rightarrow Ag$ $\mathcal{E}^\circ = -.40V$

ox $(Ag \rightarrow Ag^+ + e^-)2$ $\mathcal{E}^\circ = .40V$

$$Cd^{+2} + 2Ag \rightarrow Cd + 2Ag^+ \quad \mathcal{E}^\circ_{cell} = 1.12V$$

(b)



(c) $$\mathcal{E}_{cell} = \mathcal{E}^\circ_{cell} - \frac{RT}{nF}\ln Q$$

$$\ln \frac{>1.0}{1.0}$$

$$\mathcal{E}^\circ_{cell} - (+)(+)$$

↑ $[Ag^+]$ will ↓ the cell voltage
(↑ing the conc. of a product ↓'s the rate of the forward rxn)

(d) adding NaCl to both cells will create a ppt in the Ag cell. This will ↓ the conc of Ag^+ ion in the solution and ↓ the cell potential.

(e)

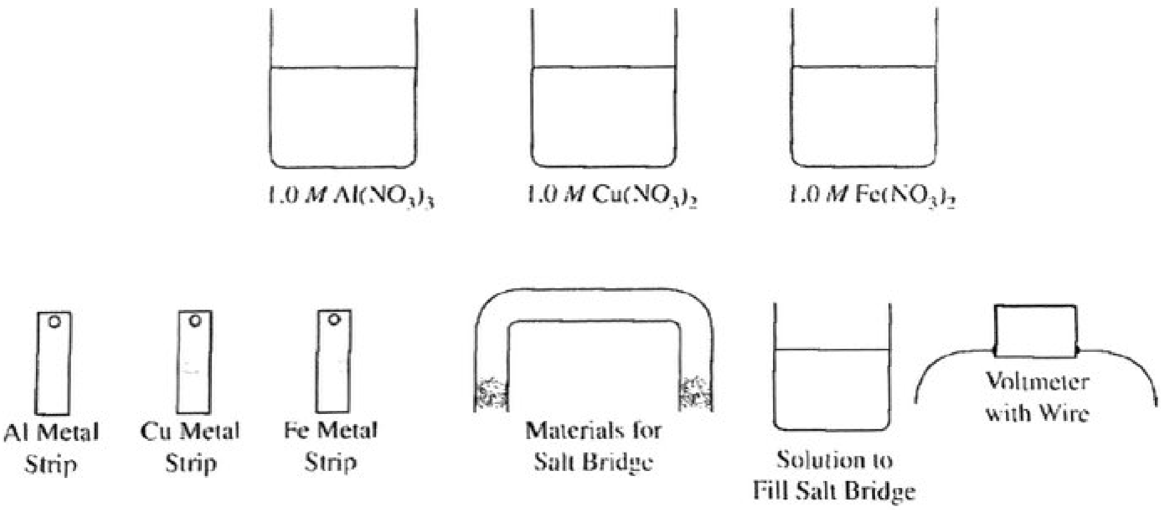
Question 6 page 4

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}.



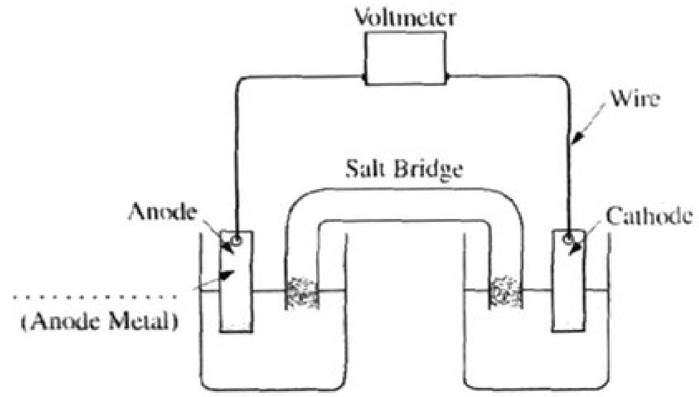
$Al(s) \rightarrow Al^{3+}(aq) + 3\,e^-$ $Cu^{2+}(aq) + 2\,e^- \rightarrow Cu(s)$ $2\,Al(s) + 3\,Cu^{2+}(aq) \rightarrow 2\,Al^{3+}(aq) + 3\,Cu(s)$	1 point for selection of correct two redox couples 1 point for correctly balanced net ionic equation

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

$Al^{3+}(aq) + 3\,e^- \rightarrow Al(s)$ $E^\circ = -1.66$ V $Cu^{2+}(aq) + 2\,e^- \rightarrow Cu(s)$ $E^\circ = +0.34$ V $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode} = +0.34\text{ V} - (-1.66\text{ V}) = +2.00\text{ V}$	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.

$E_{cell} = E^{\circ}_{cell} - 0.059 \ln\left(\frac{[Zn^{2+}]}{[Pb^{2+}]}\right)$ If $[Zn^{2+}]$ is reduced, then the ratio $\left(\frac{[Zn^{2+}]}{[Pb^{2+}]}\right) < 1$, therefore $\ln\left(\frac{[Zn^{2+}]}{[Pb^{2+}]}\right) < 0$. 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 5
page 6

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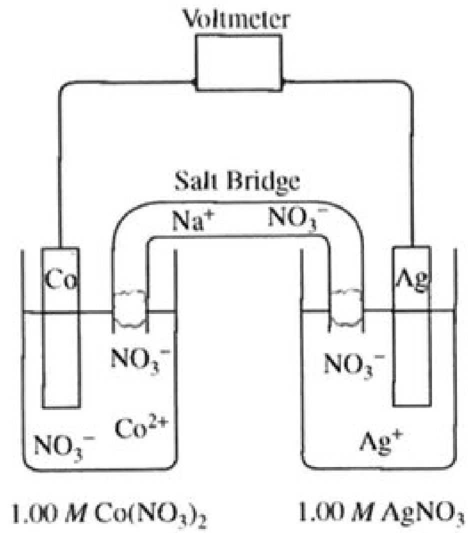
Question 5



Answer the following questions relating to the galvanic cell shown in the diagram above.

(a) Write the balanced equation for the overall cell reaction.

$2\ Ag^{+}(aq) + Co(s) \rightarrow 2\ Ag(s) + Co^{2+}(aq)$	One point is earned for the correct equation.

(b) Calculate the value of $E°$ for the cell.

$E°_{cell} = 0.80 - (-0.28) = \mathbf{1.08\ V}$	One point is earned for the correct value of $E°_{cell}$.

(c) Is the value of $\Delta G°$ for the overall cell reaction positive, negative, or 0 ? Justify your answer.

The value of $\Delta G°$ for the overall reaction must be negative because the cell reaction occurs (is spontaneous) as the cell operates. OR Since $E°_{cell}$ is positive and $\Delta G° = -nFE°$, the value of $\Delta G°$ must be negative.	One point is earned for the correct answer, including a valid justification.

Question 5 (continued)

(d) Consider the cell as it is operating.

(i) Does E_{cell} increase, decrease, or remain the same? Explain.

As the cell operates, the concentration of Ag^+ decreases and the concentration of Co^{2+} increases $\Rightarrow$ the ratio $Q = \frac{[Co^{2+}]}{[Ag^+]^2}$ increases $\Rightarrow$ $\ln Q$ increases $\Rightarrow$ $E_{cell} = E^\circ_{cell} - \frac{RT}{nF}\ln Q$ becomes smaller (decreases).	One point is earned for the correct answer, including a valid justification.

(ii) Does ΔG of the overall cell reaction increase, decrease, or remain the same? Explain.

The value of ΔG for the system increases (becomes less negative) as the cell operates and the system approaches equilibrium (when $\Delta G = 0$).	One point is earned for the correct answer, including a valid justification.

(iii) What would happen if the $NaNO_3$ solution in the salt bridge was replaced with distilled water? Explain.

The cell would not operate. The voltage of the cell is too small to overcome the electrical resistance of distilled water, which is a poor conductor due to its very low concentration of ions (about 10^{-7} *M* $H^+(aq)$ and 10^{-7} *M* $OH^-(aq)$) that could "carry" the current from one cell to the other.	One point is earned for the correct answer, including a valid justification.

(e) After a certain amount of time, the mass of the Ag electrode changes by x grams. Given that the molar mass of Ag is 108 g mol^{-1} and the molar mass of Co is 59 g mol^{-1}, write the expression for the change in the mass of the Co electrode in terms of x.

$\Delta \text{ mol Ag} = \Delta \text{ mass Ag} \times \frac{1 \text{ mol Ag}}{108 \text{ g Ag}} = x \times \frac{1}{108} = \frac{x}{108}$ $\Delta \text{ mol Co} = -\Delta \text{ mol Ag} \times \frac{1 \text{ mol Co}}{2 \text{ mol Ag}} = -\frac{x}{108} \times \frac{1}{2} = -\frac{x}{216}$ $\Delta \text{ mass Co} = \Delta \text{ mol Co} \times \frac{59 \text{ g Co}}{1 \text{ mol Co}} = -\frac{x}{216} \times 59 = -\frac{59}{216}x$	One point is earned for using the correct mole ratio of Co to Ag. One point is earned for the correct answer (negative sign is not required).