CHM 1410
Spring 2005
Test 4 (100 pts)

Name (Please Print) ______________________

$\Delta S^\circ_{rxn} = \sum nS^\circ_{(products)} - \sum mS^\circ_{(reactants)}$
$\Delta H^\circ_{rxn} = \sum n\Delta H^\circ_f \text{ (products)} - \sum n\Delta H^\circ_f \text{ (reactants)}$
$\Delta G^\circ_{rxn} = \sum n\Delta G^\circ_f \text{ (products)} - \sum n\Delta G^\circ_f \text{ (reactants)}$
$\Delta G = \Delta H - T\Delta S$ $\quad \Delta G = \Delta G^\circ + RT \ln Q$ $\quad \Delta G^\circ = -RT \ln K$
$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$ $\quad \Delta G = -nFE_{cell}$ $\quad 1F = 96{,}500 \text{ C/mol } e^-$
$\Delta G^\circ = -nFE^\circ_{cell}$ $\quad E^\circ_{cell} = RT/nF \ln K$ $\quad E = E^\circ - RT/nF \ln Q$
R = 8.314 J/mol·K $\quad E^\circ_{cell} = 0.059/n \log K$ $\quad E = E^\circ - 0.059/n \log Q$

	ΔH°_f (kJ/mol)	ΔG°_f (kJ/mol)	S° (J/K.mol)
$HBr(g)$	-36.23	-53.22	198.49
$Cl_2(g)$	0	0	222.96
$HCl(g)$	-92.30	-131.2	186.69
$Br_2(g)$	30.71	3.14	245.3

	E°_{red} (V)
$Fe^{3+} + 1e^- \rightarrow Fe^{2+}$	+0.77
$Cu^{2+} + 2e^- \rightarrow Cu$	+0.34
$Ni^{2+} + 2e^- \rightarrow Ni$	-0.25
$Cd^{2+} + 2e^- \rightarrow Cd$	-0.40

Multiple Choice (4 points each)

1. Which of the following species has the highest entropy (S°) at 25°C?

A. $CH_3OH(l)$ B. $CO(g)$ C. $MgCO_3(s)$ D. $H_2O(l)$ E. $Ni(s)$

2. A negative sign for ΔG indicates that, at constant T and P,

A. the reaction is exothermic.
B. the reaction is endothermic.
C. the reaction is fast.
D. the reaction is spontaneous.
E. ΔS must be > 0.

3. For the reaction $H_2(g) + S(s) \rightarrow H_2S(g)$, $\Delta H^\circ = -20.2$ kJ/mol and $\Delta S^\circ = +43.1$ J/K·mol. Which of the following statements is *true*?

A. The reaction is only spontaneous at low temperatures.
B. The reaction is spontaneous at all temperatures.
C. ΔG° becomes less favorable as temperature increases.
D. The reaction is spontaneous only at high temperatures.
E. The reaction is at equilibrium at 25°C under standard conditions.

4. The reaction rates of many spontaneous reactions are actually very slow. Which of
the following is the best explanation for this observation?

A. K_p for the reaction is less than one.
B. The activation energy of the reaction is large.
C. ΔG° for the reaction is positive.
D. Such reactions are endothermic.
E. The entropy change is negative.

5. Which statement is *true* for a spontaneous redox reaction carried out at standard-state conditions?

A. E°_{red} is always negative.
B. E°_{cell} is always positive.
C. E°_{ox} is always positive.
D. E°_{red} is always positive.

6. Consider the following reaction: $2Fe^{2+}(aq) + Cu^{2+} \rightarrow 2Fe^{3+}(aq) + Cu$.

When the reaction comes to equilibrium, what is the cell voltage?

A. 0.43 V B. 1.11 V C. 0.78 V D. –0.43 V E. 0 V

7. The correct statement is: _____.
A) Oxidation occurs at the anode
B) Oxidation occurs at the cathode
C) Anions travel to the cathode in an electrochemical cell.
D) Reduction occurs at the anode.

8. The unit of electromotive force is the _____.
A) erg
B) ohm
C) volt
D) ampere

9. The correct statement for a galvanic cell reaction is _____.
A) $\Delta E > 0$, $\Delta G < 0$, spontaneous
B) $\Delta E > 0$, $\Delta G > 0$, nonspontaneous
C) $\Delta E < 0$, $\Delta G < 0$, spontaneous
D) $\Delta E < 0$, $\Delta G = 0$

Problems

1. Given the following reaction:

$$2HBr(g) + Cl_2(g) \rightleftharpoons 2HCl(g) + Br_2(g)$$

(8) a) Find ΔS° for this reaction.

$\Delta S^\circ_{rxn} = \sum n S^\circ(\text{products}) - \sum m S^\circ(\text{reactants})$

$\Delta S^\circ_{rxn} = [2(186.69\ J/K\text{-mole}) + 1(245.3\ J/K\text{-mole})] - [2(198.49\ J/K\text{-mole}) + 1(222.96\ J/K\text{-mole})]$

$\Delta S^\circ_{rxn} = 618.68\ J/K - 619.94\ J/K$

$\Delta S^\circ_{rxn} = -1.26\ J/K$

(8) b) Find ΔG° for this reaction.

Either Find ΔH°_{rxn} + use $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ or

$\Delta G^\circ_{rxn} = \sum n \Delta G^\circ_f(\text{products}) - \sum m \Delta G^\circ_f(\text{reactants})$ ← Easier - less work

$\Delta G^\circ_{rxn} = [2\text{ moles}(-131.2\ kJ/mole) + 1\text{ mole}(3.14\ kJ/mole)] - [2\text{ moles}(-53.22\ kJ/mole) + 0]$

$\Delta G^\circ_{rxn} = -259.26\ kJ - (-106.44\ kJ) =$

$\Delta G^\circ_{rxn} = -152.82\ kJ$

(8) c) Find K for this reaction.

$\Delta G^\circ = -RT \ln K$

$\frac{-\Delta G^\circ}{RT} = \ln K$; $-\frac{(-152.82 \text{ kJ/mole})}{(8.314\times10^{-3} \text{ kJ/mole K})(298 K)} = \ln K$

$61.68 = \ln K$

$e^{61.68} = K$

$6.13 \times 10^{26} = K$

2.(8) Balance the following reaction: $H_2O_2 + Fe^{2+} \rightarrow Fe^{3+} + H_2O$ (acidic solution)

Oxidation: $Fe^{2+} \rightarrow Fe^{3+}$

$Fe^{2+} \rightarrow Fe^{3+} + 1e^-$

Reduction: $H_2O_2 \rightarrow H_2O$

$2e^- + 2H^+ + H_2O_2 \rightarrow 2H_2O$

$2Fe^{2+} \rightarrow 2Fe^{3+} + 2e^-$

$2e^- + 2H^+ + H_2O_2 \rightarrow 2H_2O$

$2Fe^{2+} + 2H^+ + H_2O_2 \rightarrow 2Fe^{3+} + 2H_2O$

3.(8) a) Is the following reaction spontaneous? $Cd(s) + Ni^{2+}(aq) \rightarrow Cd^{2+}(aq) + Ni(s)$
Show your work for credit.

Cathode: $Ni^{2+} + 2e^- \rightarrow Ni$

Anode: $Cd \rightarrow Cd^{2+} + 2e^-$

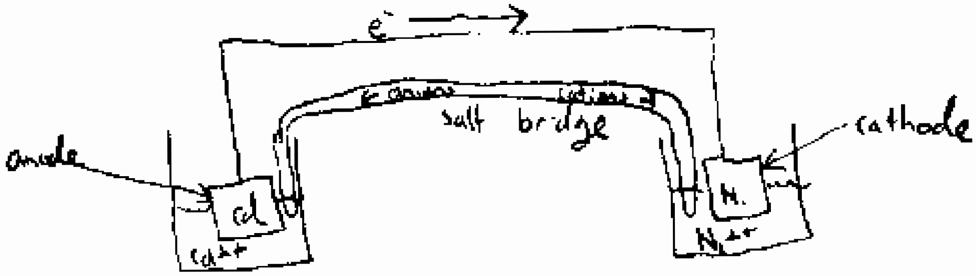
$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$E^\circ_{cell} = -0.25V - (-0.40V)$

$E^\circ_{cell} = +0.15V$

$E^\circ > 0$ so rxn is spontaneous

(8) b) Draw the basic features of a working galvanic cell based on the above reaction. Also show the cell diagram.



$Cd(s) | Cd^{2+}(aq) \,||\, Ni^{2+}(aq)\; Ni(s)$ ⇐ cell diagram

(8) c) Calculate ΔG^0 for the above reaction in this problem.

$\Delta G^\circ = -nFE^\circ$ $\quad\quad 1V = 1J/C$

$\Delta G^\circ = (-2 \text{ moles } e^-)(96{,}500 \text{ C/mole})(0.15\text{ V})$

$\Delta G^\circ = -28{,}950 \text{ J} = -28.95 \text{ kJ}$

(8) d) If the $[Cd^{2+}] = 0.1$ M, and the $[Ni^{2+}] = 0.001$ M, find E for this cell.

$E = E^\circ - \frac{0.059}{n} \log Q$

$E = +0.15V - \frac{0.059}{2} \log \frac{0.1}{0.001}$

$E = +0.15V - (0.0295)(2)$

$E = +0.15V - 0.059$

$E = 0.09$ ✓