

Please print:

Last name: _____

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Chem 1062 Exam 4

Spring 2005

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Tuesday, May 3, 2005

Equations:

$$K_{sp} = [A]^a [B]^b$$

$$\Delta S > \frac{q}{T} \text{ (spontaneous process)}$$

$$\Delta S = \frac{q}{T} \text{ (equilibrium process)}$$

$$\Delta S^\circ = \sum nS^\circ(\text{products}) - \sum mS^\circ(\text{reactants})$$

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$\Delta U = q + w$$

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

$$\Delta G^\circ = nFE^\circ_{\text{cell}}$$

$$F = 9.65 \times 10^4 \text{ C/mol e}^-$$

$$\Delta G^\circ = \sum nG_f^\circ(\text{products}) - \sum mG_f^\circ(\text{reactants})$$

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

$$R = 8.31 \text{ J/(mol}\cdot\text{K)}$$

A table of standard electrode potentials is on page 2.

Contents:

I. Multiple choice , 10 questions, 3 points each.	30 points
II. Short answer , 6 questions, worth 34 points.	34 points
III. Problems , 3 questions, worth 36 points.	<u>36 points</u>
Total: 100 points	

I, _____ have read and understand the directions given above, and pledge that I will follow all regulations with regard to Academic Dishonesty as outlined by this college when taking this exam.

Signature _____ Date and Time _____

Standard Electrode (Reduction) Potentials in Aqueous Solution at 25°C*

Cathode (Reduction) Half-Reaction	Standard Potential, $E^\circ(\text{V})$
$\text{Li}^+(aq) + e^- \rightleftharpoons \text{Li}(s)$	-3.04
$\text{Na}^+(aq) + e^- \rightleftharpoons \text{Na}(s)$	-2.71
$\text{Mg}^{2+}(aq) + 2e^- \rightleftharpoons \text{Mg}(s)$	-2.38
$\text{Al}^{3+}(aq) + 3e^- \rightleftharpoons \text{Al}(s)$	-1.66
$2\text{H}_2\text{O}(l) + 2e^- \rightleftharpoons \text{H}_2(g) + 2\text{OH}^-(aq)$	-0.83
$\text{Zn}^{2+}(aq) + 2e^- \rightleftharpoons \text{Zn}(s)$	-0.76
$\text{Cr}^{3+}(aq) + 3e^- \rightleftharpoons \text{Cr}(s)$	-0.74
$\text{Fe}^{2+}(aq) + 2e^- \rightleftharpoons \text{Fe}(s)$	-0.41
$\text{Cd}^{2+}(aq) + 2e^- \rightleftharpoons \text{Cd}(s)$	-0.40
$\text{Ni}^{2+}(aq) + 2e^- \rightleftharpoons \text{Ni}(s)$	-0.23
$\text{Sn}^{2+}(aq) + 2e^- \rightleftharpoons \text{Sn}(s)$	-0.14
$\text{Pb}^{2+}(aq) + 2e^- \rightleftharpoons \text{Pb}(s)$	-0.13
$\text{Fe}^{3+}(aq) + 3e^- \rightleftharpoons \text{Fe}(s)$	-0.04
$2\text{H}^+(aq) + 2e^- \rightleftharpoons \text{H}_2(g)$	0.00
$\text{Sn}^{4+}(aq) + 2e^- \rightleftharpoons \text{Sn}^{2+}(aq)$	0.15
$\text{Cu}^{2+}(aq) + e^- \rightleftharpoons \text{Cu}^+(aq)$	0.16
$\text{Cu}^{2+}(aq) + 2e^- \rightleftharpoons \text{Cu}(s)$	0.34
$\text{IO}^-(aq) + \text{H}_2\text{O}(l) + 2e^- \rightleftharpoons \text{I}^-(aq) + 2\text{OH}^-(aq)$	0.49
$\text{Cu}^+(aq) + e^- \rightleftharpoons \text{Cu}(s)$	0.52
$\text{I}_2(s) + 2e^- \rightleftharpoons 2\text{I}^-(aq)$	0.54
$\text{Fe}^{3+}(aq) + e^- \rightleftharpoons \text{Fe}^{2+}(aq)$	0.77
$\text{Hg}_2^{2+}(aq) + 2e^- \rightleftharpoons 2\text{Hg}(l)$	0.80
$\text{Ag}^+(aq) + e^- \rightleftharpoons \text{Ag}(s)$	0.80
$\text{Hg}^{2+}(aq) + 2e^- \rightleftharpoons \text{Hg}(l)$	0.85
$\text{ClO}^-(aq) + \text{H}_2\text{O}(l) + 2e^- \rightleftharpoons \text{Cl}^-(aq) + 2\text{OH}^-(aq)$	0.90
$2\text{Hg}_2^{2+}(aq) + 2e^- \rightleftharpoons \text{Hg}_2^{2+}(aq)$	0.90
$\text{NO}_3^-(aq) + 4\text{H}^+(aq) + 3e^- \rightleftharpoons \text{NO}(g) + 2\text{H}_2\text{O}(l)$	0.96
$\text{Br}_2(l) + 2e^- \rightleftharpoons 2\text{Br}^-(aq)$	1.07
$\text{O}_2(g) + 4\text{H}^+(aq) + 4e^- \rightleftharpoons 2\text{H}_2\text{O}(l)$	1.23
$\text{Cr}_2\text{O}_7^{2-}(aq) + 14\text{H}^+(aq) + 6e^- \rightleftharpoons 2\text{Cr}^{3+}(aq) + 7\text{H}_2\text{O}(l)$	1.33
$\text{Cl}_2(g) + 2e^- \rightleftharpoons 2\text{Cl}^-(aq)$	1.36
$\text{MnO}_4^-(aq) + 8\text{H}^+(aq) + 5e^- \rightleftharpoons \text{Mn}^{2+}(aq) + 4\text{H}_2\text{O}(l)$	1.49
$\text{H}_2\text{O}_2(aq) + 2\text{H}^+(aq) + 2e^- \rightleftharpoons 2\text{H}_2\text{O}(l)$	1.78
$\text{S}_2\text{O}_8^{2-}(aq) + 2e^- \rightleftharpoons 2\text{SO}_4^{2-}(aq)$	2.01
$\text{F}_2(g) + 2e^- \rightleftharpoons 2\text{F}^-(aq)$	2.87

I. Multiple choice

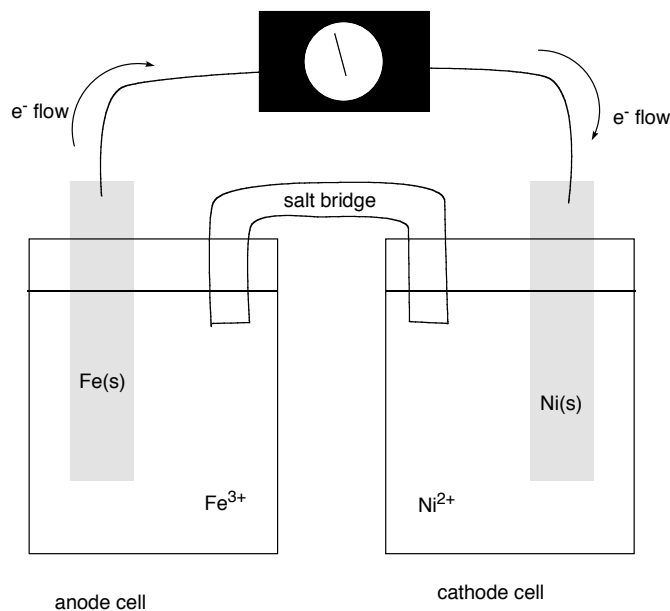
Choose the best answer from the choices given, and clearly **circle** the letter of your choice. You may need to consult the table on page 2 for some answers. (3 pts each)

- What is the solubility product expression for $\text{Au}(\text{OH})_3$?
 - $K_{sp} = [\text{Au}^{3+}][\text{OH}^-]$
 - $K_{sp} = [\text{Au}^{3+}][\text{OH}^-]^3$
 - $K_{sp} = [\text{Au}^{3+}][3\text{OH}^-]$
 - $K_{sp} = [\text{Au}^{3+}][3\text{OH}^-]^3$
 - $K_{sp} = 3[\text{Au}^{3+}][\text{OH}^-]^3$
- What is the solubility product (K_{sp}) for SrF_2 if its molar solubility is $1.0 \times 10^{-3} \text{ M}$?
 - 1.0×10^{-9}
 - 2.0×10^{-9}
 - 4.0×10^{-9}
 - 1.0×10^{-3}
 - 2.0×10^{-3}
- The total entropy of a system and its surroundings always increases for a spontaneous process. This is a statement of
 - the law of constant composition.
 - the first law of thermodynamics.
 - the second law of thermodynamics.
 - the third law of thermodynamics.
 - the law of conservation of matter.
- For a spontaneous process, which of the following is **always** negative?
 - ΔH
 - ΔS
 - ΔG
 - K
 - $T\Delta S$
- Which of the following statements is **false**?
 - ΔH° can be assumed to be relatively temperature independent.
 - ΔS° can be assumed to be relatively temperature independent.
 - ΔG° can be assumed to be relatively temperature independent.
 - K is temperature dependent
 - If $\Delta G^\circ = 0$, then $K = 1$.

6. If a nonspontaneous reaction has a negative ΔH° value, which of the following statements is **true**?
- ΔS° is positive, and the reaction may become spontaneous at higher temperatures.
 - ΔS° is positive, and the reaction may become spontaneous at lower temperatures.
 - ΔS° is negative, and the reaction may become spontaneous at higher temperatures.
 - ΔS° is negative, and the reaction may become spontaneous at lower temperatures.
 - ΔS° is negative, and the reaction is nonspontaneous at all temperatures.
7. For a reaction that has an equilibrium constant of 3×10^5 , which of the following statements must be **true**?
- ΔG° is positive
 - ΔG° is negative
 - ΔH° is positive
 - ΔH° is negative
 - ΔS° is positive
8. For the voltaic cell represented by $\text{Fe}(s) | \text{Fe}^{2+}(aq) || \text{Ag}^+(aq) | \text{Ag}(s)$, which of the following statements is **true**?
- The cathode half-reaction is $\text{Fe}(s) \rightarrow \text{Fe}^{2+}(aq) + 2e^-$
 - The anode half-reaction $\text{Ag}^+(aq) + e^- \rightarrow \text{Ag}(s)$
 - $\text{Fe}(s)$ is oxidized
 - none of these are true
 - all of these are true
9. Which of the following voltaic cell reactions will produce the highest voltage?
- $2\text{H}^+(aq) + \text{Zn}(s) \rightarrow \text{H}_2(g) + \text{Zn}^{2+}(aq)$
 - $\text{Ni}^{2+}(aq) + \text{Cd}(s) \rightarrow \text{Ni}(s) + \text{Cd}^{2+}(aq)$
 - $\text{Sn}^{2+}(aq) + \text{Mg}(s) \rightarrow \text{Sn}(s) + \text{Mg}^{2+}(aq)$
 - $\text{Hg}_2^{2+}(aq) + \text{Cu}(s) \rightarrow 2\text{Hg}(l) + \text{Cu}^{2+}(aq)$
 - $\text{Cu}^{2+}(aq) + \text{Sn}(s) \rightarrow \text{Cu}(s) + \text{Sn}^{2+}(aq)$
10. If a beta particle is ejected from a nucleus, the following occurs:
- a positron is converted to a neutron.
 - a neutron is converted to a positron.
 - a neutron is converted to a proton.
 - a proton is converted to a neutron.
 - a neutron is converted to an electron.

II. Conceptual

11. Sketch the galvanic cell represented by $\text{Fe}(s) | \text{Fe}^{3+}(aq) || \text{Ni}^{2+}(aq) | \text{Ni}(s)$. Label the following: anode cell, cathode cell, salt bridge, $\text{Fe}(s)$, $\text{Fe}^{3+}(aq)$, $\text{Ni}^{2+}(aq)$, $\text{Ni}(s)$. Also draw a voltmeter or light bulb connected to the two electrodes, and indicate the direction of electron flow in the wires connected to the electrodes. (10 pts)



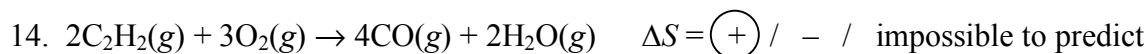
For the following reactions, predict and **circle** the sign of ΔS , and explain why. If ΔS is impossible to predict, explain why as well. (4 pts each)



Rationale: 3 moles of gas are converted into 2 moles of gas, 3 molecules are combined into one



Rationale: moles of gas are the same, phases are the same

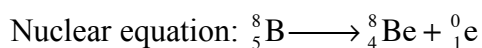


Rationale: 5 moles of gas are converted into 6 moles of gas

For the following nuclei, predict and **circle** the type of radioactive decay they will undergo based upon their N/Z ratio, and write a balanced nuclear equation for the decay. (6 pts each)

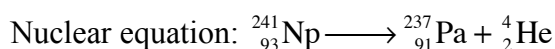
15. ${}^8_5\text{B}$ N/Z ratio: 3/5

Radioactive decay: alpha emission / beta emission / positron emission



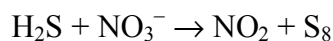
16. ${}^{241}_{93}\text{Np}$ N/Z ratio: 148/93

Radioactive decay: alpha emission / beta emission / positron emission

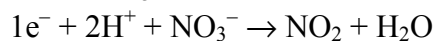
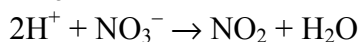
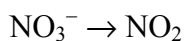
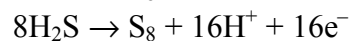
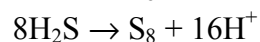
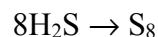
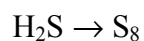


III. Problems

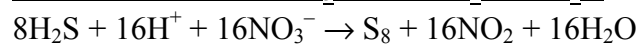
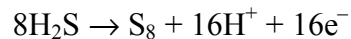
17. (10 pts) Consider the following unbalanced redox reaction, which takes place in an **acidic** solution:



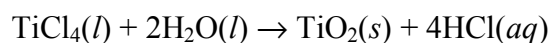
- a. Write the two half-reactions with all atoms and charges balanced. Show your work for each step.



- b. Using the half-reactions above, write the total balanced redox equation. Show your work.



18. (16 pts) Consider the following reaction:



- a. Using the table of thermodynamic values given to the right, calculate ΔH° (in kJ) and ΔS° (in J/K) for this reaction.

Substance	ΔH_f° (kJ/mol)	S° (J/mol·K)
$\text{TiCl}_4(l)$	-804.2	221.9
$\text{H}_2\text{O}(l)$	-285.83	69.91
$\text{TiO}_2(s)$	-944.7	50.29
$\text{HCl}(aq)$	-167.2	56.5

$$\begin{aligned}\Delta H^\circ &= [-944.7 + 4(-167.2)] - [-804.2 + 2(-285.83)] \\ &= -237.64 \text{ kJ}\end{aligned}$$

$$\begin{aligned}\Delta S^\circ &= [50.29 + 4(56.5)] - [221.9 + 2(69.91)] \\ &= -85.43 \text{ J/K}\end{aligned}$$

- b. Calculate ΔG° (in kJ) for this reaction at 298 K. Is this reaction spontaneous at room temperature?

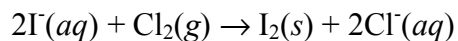
$$\begin{aligned}\Delta G^\circ &= \Delta H^\circ - T \Delta S^\circ \\ &= -237.64 \text{ kJ} - (298 \text{ K})(-85.43 \times 10^{-3} \text{ kJ/K}) \\ &= -212.18 \text{ kJ (the reaction is spontaneous at room temperature)}\end{aligned}$$

- c. Calculate the equilibrium constant K at 298 K. Explain how this value is consistent with the value of ΔG° you obtained.

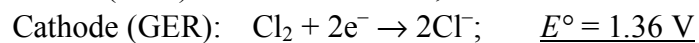
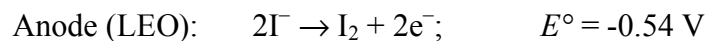
$$\begin{aligned}\Delta G^\circ &= -RT \ln K \\ -212.18 \times 10^3 \text{ J} &= -(8.31 \text{ J/mol·K})(298 \text{ K}) \ln K \\ \ln K &= 85.68 \\ K &= 1.63 \times 10^{37}\end{aligned}$$

(this very large K value corresponds with a negative ΔG° because $K > 1$ when $\Delta G^\circ < 0$.)

19. (10 pts) Consider the following balanced redox reaction:



a. Using the values in the table on page 2, calculate E_{cell} for this reaction.



$$E_{\text{CELL}} = 0.82 \text{ V}$$

$$\text{Or simply } E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}} = 1.36 - 0.54 = 0.82 \text{ V}$$

b. Calculate ΔG° for this reaction. Is it spontaneous at room temperature?

$$\square \Delta G^\circ = -nFE_{\text{cell}} = -(2 \text{ mol e}^-)(9.65 \times 10^4 \text{ C/mol e}^-)(0.82 \text{ V})$$

$$= -158,260 \text{ C} \cdot \text{V} = -158,260 \text{ J} = -160 \text{ kJ (the reaction is spontaneous)}$$

-END OF EXAM-