Chemical Reactivity in the Environment
Final Exam - practice

Read all questions carefully. Show all work for full credit. Work independently, using only a calculator (with no pre-programmed information relevant to this course) and a periodic table to assist you. If the meaning of a question is unclear, ask the instructor.

SECTION I – Basics of chemical reactivity, introduction to thermodynamics

1. The second law of thermodynamics states:
   a. The change in internal energy of a system is equal to the work done by the system and the heat flowing out of the system.
   b. The change in the entropy of the universe is always greater than zero.
   c. Energy is equal to mass times the speed of light squared.

2. The amount of work that can be done by a chemical system is represented by:
   a. ΔG
   b. ΔG°
   c. the heat released when the reaction occurs
   d. RTlogK

3. The Haber process is the human equivalent to biological nitrogen fixation, except that it must occur under relatively extreme conditions. It is generally carried out at a pressure of 250 atm and at a temperature of 400° C. What is the free energy change associated with this reaction? (Assume initially that the reactants each have a partial pressure of 125 atm and that NH₃ has a partial pressure of 1 x 10⁻⁴ atm.) Answer on answer sheet should be the ΔG for the reaction.
4. Dinitrogen pentoxide decomposes to form NO₂ + O₂ + NO. If 0.03 moles of N₂O₅ reacted in 500 mL of water, by how much would the temperature of the water change? Be sure to include if the temperature of the water would go up or down. Answer on the answer sheet should be +/- ___K

5. Given the table of bond energies below, calculate the enthalpy change for the combustion of one mole of methane (CH₄) in air. Answer on the answer sheet should be ΔH as calculated by this method.

<table>
<thead>
<tr>
<th>Bond</th>
<th>Energy (kJ mol⁻¹)</th>
<th>Bond</th>
<th>Energy (kJ mol⁻¹)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C-H</td>
<td>414</td>
<td>C=O</td>
<td>799</td>
</tr>
<tr>
<td>O=O</td>
<td>498</td>
<td>C-O</td>
<td>360</td>
</tr>
<tr>
<td>O-O</td>
<td>142</td>
<td>O-H</td>
<td>464</td>
</tr>
</tbody>
</table>

SECTION II – Equilibrium, Bronsted Acid/base chemistry

6. At low pH, a weak base will be mostly ionized/mostly not ionized (select one and put on answer sheet)
7. Q represents:
   a. the concentration of all species at equilibrium
   b. the energy stored in the system because the system is not at equilibrium
   c. the heat stored in the system because the system is not at equilibrium
   d. a measure of the disorder of the system

8. Calculate the pH of a solution with 0.1 moles H$_2$CO$_3$ and 0.25 moles HCO$_3^-$ in 200 mL. The $K_a$ for carbonic acid is $4.3 \times 10^{-7}$. Answer sheet should give pH only.

9. What is the pH, pOH, [H$^+$], and [OH$^-$] of the following solution?

   6 ml 0.7 M HNO$_3$ (aq)
   12 ml 0.1M Ca(OH)$_2$ (aq)
   22 mL of water

   (total volume of the solution is 40 mL) Answer sheet should give pH, pOH, [H$^+$] and [OH$^-$]

10. Calculate the concentration of all species at equilibrium give the initial concentrations and K for the following reaction:

    H$_2$(g) + I$_2$(g) $\rightarrow$ HI(g)
K = 50.5
Initial [ ]:
HI = 2.0 M
I₂ = 3.0 M
H₂ = 5.0 M

SECTION III – Solubility of ionic solids, redox reactions

11. For the following reaction, as pH becomes larger, would the cell potential (E) increase or decrease? Write either “increase” or “decrease” on the answer sheet.

Fe²⁺ + O₂ → Fe³⁺ + H₂O

12. As pH becomes smaller would the solubility of Fe(OH)₃(s) decrease or increase? Write either “increase” or “decrease” on the answer sheet.
13. How many grams of Al(OH)$_3$(s) will saturate 50.0 mL of water at 25 °C given that the $K_{sp}$ for Aluminum hydroxide is $1.9 \times 10^{-33}$?

14. Balance the following redox reaction: NaBr(s) + SO$_4^{2-}$(aq) $\rightarrow$ Br$_2$(l) + SO$_2$ (g) + Na$^+$ (aq)

15. Find $\Delta G^\circ$ and $E^\circ$ for the following reaction:

$I_2(s) + Al(s) \rightarrow I^-(aq) + Al^{3+}(aq)$. Write $\Delta G^\circ$ and $E^\circ$ on the answer sheet.

$I_2 + 2e^- \rightarrow 2I^-$ $E^\circ = 0.54$
$Al^{3+} + 3e^- \rightarrow Al(s)$ $E^\circ = -1.66V$

SECTION V – Integrated Problems, Lab, problems just for fun.

23. Calculate the pH of rainwater in equilibrium with 0.12 ppm SO$_2$ in the troposphere. $K_H = 1.2 \text{ M atm}^{-1}$, SO$_2$ + H$_2$O $\rightarrow$ H$_2$SO$_3$. H$_2$SO$_3$ has a $K_a = 1.7 \times 10^{-2}$. 
24. How many grams of Ag\(^+\) (from AgCl) will be found dissolved in 50 mL of a 0.1 M solution of AgNO\(_3\) given that the Ksp for AgCl is 1.8 x 10\(^{-8}\)? What is the \(\Delta G^o\) for the dissolution of AgCl?

25. Describe how you would make up a 100 ppm nitrate solution from KNO\(_3\) \cdot 2H\(_2\)O. Describe how you would use this stock solution to make up 50 mL of a 1 ppm nitrate solution.

26. Given a solution of [Fe(phen)\(_3\)]\(^{2+}\) with an absorbance of 0.5 and an extinction coefficient of 11.1 mM\(^{-1}\)cm\(^{-1}\), calculate the concentration of [Fe(phen)\(_3\)]\(^{2+}\). The path length is 1 cm.

27. Matching – Match letter and number of correct name and chemical formula. Put on answer key

   a) nitrate
   1.NO\(_2^−\)
| b) nitrite          | 2. NH₃          |
| c) carbonate       | 3. OH⁻          |
| d) phosphate       | 4. HCO₃⁻        |
| e) hydroxide       | 5. H₂S          |
| f) ammonium        | 6. SO₃²⁻        |
| g) hydrogen sulfide| 7. NO₃⁻         |
| h) sulfite         | 8. SO₄²⁻        |
| i) carbon dioxide  | 9. CO₂          |
| j) dimethylsulfide | 10. (CH₃)₂S     |
|                     | 11. S²⁻         |
|                     | 12. CO₃²⁻       |
|                     | 13. PO₄³⁻       |
|                     | 14. NH₄⁺        |
|                     | 15. PO₂²⁻       |

28. Describe what happens when you titrate a strong acid and a strong base. Describe what happens when you titrate a weak acid and its conjugate weak base.

**Numbers and Formulas you may need:**

<table>
<thead>
<tr>
<th></th>
<th>Δ Gᵣ (kJ)</th>
<th>Δ Hᵣ (kJ)</th>
<th>S (J K⁻¹)</th>
</tr>
</thead>
<tbody>
<tr>
<td>NH₃(g)</td>
<td>-17</td>
<td>-46</td>
<td>193</td>
</tr>
<tr>
<td>NH₃(aq)</td>
<td>-27</td>
<td>-80</td>
<td>111</td>
</tr>
<tr>
<td>NH₄⁺(aq)</td>
<td>-79</td>
<td>-132</td>
<td>113</td>
</tr>
<tr>
<td>NO(g)</td>
<td>87</td>
<td>90</td>
<td>211</td>
</tr>
<tr>
<td>NO₂(g)</td>
<td>52</td>
<td>34</td>
<td>240</td>
</tr>
<tr>
<td>N₂O(g)</td>
<td>104</td>
<td>82</td>
<td>220</td>
</tr>
<tr>
<td>N₂O₄(g)</td>
<td>98</td>
<td>10</td>
<td>304</td>
</tr>
</tbody>
</table>
\[ F = 96,485 \, \text{C mol}^{-1} \]
\[ \Delta G = \Delta G^0 + RT \ln Q \]
\[ R = 8.314 \, \text{J K}^{-1} \text{mol}^{-1} \]
\[ 0 \, \text{K} = -273^\circ \text{C} \]
\[ \text{Rate} = k \]
\[ \text{Rate} = k[A] \]
\[ \text{Rate} = k[A]^2 \]
\[ [A] = -kt + [A]_0 \] (0 order)
\[ \ln[A] = -kt + \ln[A]_0 \]
\[ t_{1/2} = \ln 2/k \]
\[ R = 0.08206 \, \text{L atm K}^{-1} \text{mol}^{-1} \]
\[ \Delta G^0 = -RT \ln K \]
\[ K_p = K_c (RT)^{\Delta n} \]
\[ \Delta G = \Delta G^0 + RT \ln Q \]

\[ E = E^0 - 0.0591 \frac{\ln Q}{n} \] at 298 K

\[ K_{H(CO_2)} = 3.4 \times 10^{-2} \, \text{M atm}^{-1} \]
\[ \text{Fe(OH)}_2 \quad K_{sp} = 1.8 \times 10^{-15} \]
\[ \text{Fe(OH)}_3 \quad K_{sp} = 4.0 \times 10^{-38} \]
Fe$^{3+} + e^- \rightarrow Fe^{2+} \quad E_{\text{red}}^o = .77 \ \text{V}$

$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O \quad E_{\text{red}}^o = 1.23 \ \text{V}$

$S + 2e^- + 2H^+ \rightarrow S^{2-} \quad E_{\text{red}}^o = .123V$

$\Delta H = nC_v\Delta T \text{ or } gC_p\Delta T$

$\Delta G = -nFE$

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

$A = bce$

$C_p = 4.184 \ \text{J} \ \text{g}^{-1}\text{K}^{-1}$