Chem. 1B Midterm 1  
Version B  
Jan. 29, 2016

Name:________________________________________
Print Neatly. You will lose 1 point if I cannot read your name or perm number.

Perm Number:__________________________________

All work must be shown on the exam for partial credit. Points will be taken off for incorrect or no units. Calculators are allowed. Cell phones may not be used for calculators. On fundamental and short answer problems you must show your work in order to receive credit for the problem. **If your cell phone goes off during the exam you will have your exam removed from you.**

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Fundamental Questions
Each of these fundamental chemistry questions is worth 6 points. You must show work to get credit. Little to no partial credit will be rewarded. Make sure to include the correct units on your answers.

1) 6 pts  A jelly donut (55 g) is burning in a bomb calorimeter resulting in a temperature rise of 25.0°C. The heat capacity of the calorimeter is 41.7 kJ/°C. How many calories are in the jelly donut?

\[
\Delta E_{com} = -q_{cal} = -C_{cal}\Delta T = - \left(41.7 \frac{kJ}{\degree C}\right) (25.0 \degree C)
\]
\[
\Delta E_{com} = -1040 \text{kJ} \left(\frac{1000 \text{J}}{1 \text{kJ}}\right) \left(\frac{1 \text{cal}}{4.184 \text{J}}\right) = -249000 \text{ cal}
\]
or -249 Cal (kcal)

2) 6 pts  Given the following data:

\[
P_4(s) + 6\text{Cl}_2(g) \rightarrow 4\text{PCl}_3(g) \quad \Delta H = 1225.6 \text{ kJ}
\]
\[
P_4(s) + 5\text{O}_2(g) \rightarrow \text{P}_4\text{O}_{10}(s) \quad \Delta H = -2967.3 \text{ kJ}
\]
\[
\text{PCl}_3(g) + \text{Cl}_2(g) \rightarrow \text{PCl}_5(g) \quad \Delta H = -84.2 \text{ kJ}
\]
\[
\text{PCl}_3(g) + \frac{1}{2}\text{O}_2(g) \rightarrow \text{Cl}_3\text{PO}(g) \quad \Delta H = -285.7 \text{ kJ}
\]

Calculate \(\Delta H\) for the reaction

\[
P_4\text{O}_{10}(s) + 6\text{PCl}_5(g) \rightarrow 10\text{Cl}_3\text{PO}(g)
\]

\[
\Delta H = \Delta H_a + \Delta H_b + \Delta H_c + \Delta H_d
\]
\[
\Delta H = 1225.6 \text{ kJ} + 2967.3 \text{ kJ} + 505.2 \text{ kJ} + -2857 \text{ kJ} = -610.1 \text{ kJ}
\]

3) 6 pts  What is \(K\) at 25°C for the reaction \(2\text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g)\).

\[
\Delta G^\circ = -RT\ln (K)
\]

Determine \(\Delta G^\circ_{rxn}\)

\[
\Delta G^\circ_{rxn} = \sum \Delta G^\circ_f (prod) - \sum \Delta G^\circ_f (react)
\]
\[
\Delta G^\circ_{rxn} = 2\Delta G^\circ_f (\text{SO}_3(g)) - 2\Delta G^\circ_f (\text{SO}_2(g)) - \Delta G^\circ_f (\text{O}_2(g))
\]
\[
\Delta G^\circ_{rxn} = 2 \left(-371 \frac{kJ}{mol}\right) - 2 \left(-300 \frac{kJ}{mol}\right) - 0 \frac{kJ}{mol} = -142 \frac{kJ}{mol}
\]

\[
K = e^{\frac{-\Delta G^\circ_{rxn}}{RT}} = e^{\frac{-142 \frac{kJ}{mol}}{0.0083145 \frac{kJ}{mol K}(298 K)}} = 7.76 \times 10^{24}
\]
4a)  3 pts. For the reaction:  \(2\text{O}(g) \rightarrow \text{O}_2(g)\)  
Predict the signs of \(\Delta H\) and \(\Delta S\).

\(\Delta H = -\) (heat must go on side with \(\text{O}_2\) because heat is needed to break bonds)
\(\Delta S = -\) (possitional probably decreases when going from 2 to 1)

4b)  3 pts. Would the reaction be more spontaneous at high or low temperatures?

\[\Delta G = \Delta H - T\Delta S\]  
Therefore, the reaction is spontaneous at low temp

5)  6 pts Calculate \(\Delta G^\circ\) at 125\(^\circ\)C for the following reaction. Assume that \(\Delta H^\circ\) and \(\Delta S^\circ\) are temperature independent.

\(2\text{CH}_3\text{OH}(g) + 3\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 4\text{H}_2\text{O}(l)\)

\[\Delta G_{\text{rxn}}^\circ = \Delta H_{\text{rxn}}^\circ - T\Delta S_{\text{rxn}}^\circ\]

\[\Delta H_{\text{rxn}}^\circ = \sum \Delta H_f^\circ (\text{prod}) - \sum \Delta H_f^\circ (\text{react})\]

\[\Delta H_{\text{rxn}}^\circ = 2\Delta H_f^\circ (\text{CO}_2) + 4\Delta H_f^\circ (\text{H}_2\text{O}) - 2\Delta H_f^\circ (\text{CH}_3\text{OH}) - 3\Delta H_f^\circ (\text{O}_2)\]

\[\Delta H_{\text{rxn}}^\circ = 2(-393.5 \frac{\text{kJ}}{\text{mol}}) + 4(-286 \frac{\text{kJ}}{\text{mol}}) - 2(-201 \frac{\text{kJ}}{\text{mol}}) - 3(0 \frac{\text{kJ}}{\text{mol}}) = -1529 \frac{\text{kJ}}{\text{mol}}\]

\[\Delta S_{\text{rxn}}^\circ = \sum S_{\text{prod}}^\circ - \sum S_{\text{react}}^\circ\]

\[\Delta S_{\text{rxn}}^\circ = 2S_{\text{CO}_2}^\circ + 4S_{\text{H}_2\text{O}}^\circ - 2S_{\text{CH}_3\text{OH}}^\circ - 3S_{\text{O}_2}^\circ\]

\[\Delta S_{\text{rxn}}^\circ = 2(214 \frac{\text{J}}{\text{mol} \cdot \text{K}}) + 4(70.7 \frac{\text{J}}{\text{mol} \cdot \text{K}}) - 2(240 \frac{\text{J}}{\text{mol} \cdot \text{K}}) - 3(205 \frac{\text{J}}{\text{mol} \cdot \text{K}}) = -387 \frac{\text{J}}{\text{mol} \cdot \text{K}}\]

\[\Delta G_{\text{rxn}}^\circ = \Delta H_{\text{rxn}}^\circ - T\Delta S_{\text{rxn}}^\circ\]

\[\Delta G_{\text{rxn}}^\circ = -1529 \frac{\text{kJ}}{\text{mol}} - (398k) (-0.387 \frac{\text{kJ}}{\text{mol} \cdot \text{K}}) = -1375 \frac{\text{kJ}}{\text{mol}}\]

6)  6 pts A chemist fills a reaction vessel with 9.34 atm methane (\(\text{CH}_4\)) gas, 1.58 atm oxygen (\(\text{O}_2\)) gas, 4.28 atm carbon dioxide (\(\text{CO}_2\)) gas, and 7.43 atm hydrogen (\(\text{H}_2\)) gas at a temperature of 25.0\(^\circ\)C. Under these conditions, calculate the reaction free energy \(\Delta G\) for the following chemical reaction:

\(\text{CH}_4(g) + \text{O}_2(g) \rightleftharpoons \text{CO}_2(g) + 2\text{H}_2(g)\)

Round your answer to the nearest kilojoule.

\[\Delta G = \Delta G^\circ + RT\ln(Q) = \Delta G^\circ + RT\ln \left(\frac{P_{\text{CO}_2}P_{\text{H}_2}^2}{P_{\text{CH}_4}P_{\text{O}_2}}\right)\]

\[\Delta G_{\text{rxn}}^\circ = \Delta G_f^\circ (\text{CO}_2) + 2\Delta G_f^\circ (\text{H}_2) - \Delta G_f^\circ (\text{CH}_4) - \Delta G_f^\circ (\text{O}_2)\]

\[\Delta G_{\text{rxn}}^\circ = \left(-394 \frac{\text{kJ}}{\text{mol}}\right) + 2\left(0 \frac{\text{kJ}}{\text{mol}}\right) - \left(-51 \frac{\text{kJ}}{\text{mol}}\right) - \left(0 \frac{\text{kJ}}{\text{mol}}\right) = -343 \frac{\text{kJ}}{\text{mol}}\]

\[\Delta G = -343 \frac{\text{kJ}}{\text{mol}} + \left(0.0083145 \frac{\text{kJ}}{\text{mol} \cdot \text{K}}\right)(298.2\text{K})\ln \left(\frac{(4.28)^2(7.43)}{(9.34)(1.58)}\right)\]

\[= -336 \frac{\text{kJ}}{\text{mol}}\]
Short Answer Questions

Each of the following short answer questions are worth the noted points. Partial credit will be given. You must show your work to get credit. Make sure to include proper units on your answer.

1a) 12 pts  You are camping and want to make yourself a cup of coffee in the morning with your butane (C₅H₁₀) stove. Before you go to bed you set out 250 mL of water to use for your coffee. When you get up in the morning your water has frozen, you use a thermometer and find the temperature of the ice to be 0°C. How much butane, in grams, will you use to heat your water to your ideal coffee temperature of 75°C. Assume no heat from the combustion reaction is lost to the surroundings and that H₂O(g) forms in the reaction. The ΔH_fus(water), ΔH_vap(water), ΔH_f(C₅H₁₀) are 6.01 \( \frac{kJ}{mol} \), 40.65 \( \frac{kJ}{mol} \), and -126 \( \frac{kJ}{mol} \) respectively. The density of water is 1.00 \( \frac{g}{mL} \).

Reaction of Interest

\[ 2C₅H₁₀(g) + 13O₂(g) \rightarrow 8CO₂(g) + 10H₂O(g) \]

Determine \( \Delta H^\circ_{rxn} \)

\[ \Delta H^\circ_{rxn} = 8\Delta H^\circ_f(CO₂) + 10\Delta H^\circ_f(H₂O) - 2\Delta H^\circ_f(C₅H₁₀) - 13\Delta H^\circ_f(O₂) \]

\[ \Delta H^\circ_{rxn} = 8(-393.5 \frac{kJ}{mol}) + 10(-242 \frac{kJ}{mol}) - 2(-126 \frac{kJ}{mol}) - 13(0 \frac{kJ}{mol}) = -5316 \frac{kJ}{mol} \]

Reaction takes place at constant pressure therefore \( \Delta H_{rxn} = q_{rxn} \)

All the heat from the reaction goes into the ice/water

\[ q_{rxn} = -q_{ice/water} = -(q_{melting} + q_{water}) \]

\[ q_{melting} = n_{H₂O} \Delta H_{fus} \]

\[ n_{H₂O} = 250 mL \left( \frac{1.00 g}{1 mL} \right) \left( \frac{1 mol}{18.0148 g} \right) = 13.9 mol \]

\[ q_{melting} = (13.9 mol) \left( 601 \frac{kJ}{mol} \right) = 83.5 kJ \]

\[ q_{water} = m_{H₂O} C_{H₂O} \Delta T = (250 g) \left( 4.184 \frac{J}{g°C} \right) (75°C - 0°C) = 78500 J \]

\[ = 78.5 kJ \]

\[ q_{ice/water} = 83.5 kJ + 78.5 kJ = 162.0 kJ = -q_{rxn} \]

\[ -162.0 kJ \left( \frac{2 mol C₅H₁₀}{-5316 kJ} \right) \left( \frac{58.123 g}{1 mol C₅H₁₀} \right) = 3.542 g \]

1b) 8 pts  What is ΔE, ΔH, and ΔS for the ice/water that was heated.

The ice and the water do not change size therefore no work is done and

\[ \Delta E = q = 162.0 kJ \]

The process takes place at constant pressure

\[ \Delta H = q = 162.0 kJ \]

To calculate ΔS break into 2 parts \( \Delta S_{melting} \) and \( \Delta S_{water} \)

\[ \Delta S_{melting} = \frac{\Delta H_{melting}}{T} = \frac{83.5 kJ}{273 K} = 0.306 \frac{kJ}{K} = 306 \frac{J}{K} \]

\[ \Delta S_{water} = m_{H₂O} C_{H₂O} ln \left( \frac{T_2}{T_1} \right) = (250 g) \left( 4.184 \frac{J}{g°C} \right) ln \left( \frac{348 K}{273 K} \right) = 254 \frac{J}{K} \]

\[ \Delta S_{tot} = 306 \frac{J}{K} + 254 \frac{J}{K} = 560 \frac{J}{K} \]
2) **14 pts** 350. mL of 7.0 M H₂CO₃ and 300. mL of 15.0 M NaOH are mixed in a constant pressure calorimeter with heat capacity of 54 J/°C. The initial temperature of both solutions are 25.00°C and the final temperature of the system is 24.84°C. What is ΔHₚᵣₓ in kJ/mol. Assume that the heat capacity of the solution is 4.184 J/°C and the density of the solution is 1.00 g/mL.

Reaction of Interest

\[ \text{H}_2\text{CO}_3(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{Na}_2\text{CO}_3(\text{aq}) + 2\text{H}_2\text{O}(\text{l}) \]

\[ \Delta H_{\text{rxn}} = q_{\text{rxn}} = -(q_{\text{sol}} + q_{\text{cal}}) \]

\[ q_{\text{sol}} = m_{\text{sol}}c_{\text{sol}}\Delta T = (650.\, g) \left(4.184 \frac{J}{g \cdot ^\circ C}\right)(24.84^\circ C - 25.00^\circ C) \]

\[ = -435\, J \]

\[ q_{\text{cal}} = C_{\text{cal}}\Delta T = \left(54 \frac{J}{^\circ C}\right)(24.84^\circ C - 25.00^\circ C) = -8.6\, J \]

\[ \Delta H_{\text{rxn}} = -(435\, J + -8.6\, J) = 443\, J \]

Change to \( \frac{kJ}{mol} \)

This is a limiting reagent problem

<table>
<thead>
<tr>
<th></th>
<th>H₂CO₃</th>
<th>NaOH (L.R.)</th>
<th>Na₂CO₃</th>
<th>H₂O</th>
</tr>
</thead>
<tbody>
<tr>
<td>l(mol)</td>
<td>0.350L(7.0M)</td>
<td>0.300L(15.0M)</td>
<td>0 mol</td>
<td>N/A</td>
</tr>
<tr>
<td></td>
<td>= 2.45 mol</td>
<td>= 4.50 mol</td>
<td></td>
<td></td>
</tr>
<tr>
<td>F(mol)</td>
<td>2.45-2.25=0.20 mol</td>
<td>0 mol</td>
<td>2.25 mol</td>
<td>N/A</td>
</tr>
</tbody>
</table>

\[ \Delta H_{\text{rxn}} = \frac{443\, J}{4.50\, mol} = 98.4\, \frac{J}{mol} \]

This is for NaOH which has coefficient of 2 in front of it, need to convert to something with a coefficient of 1 in front (ex: H₂CO₃) therefore multiply by 2.

\[ \Delta H_{\text{rxn}} = 197\, \frac{J}{mol} = 0.197\, \frac{kJ}{mol} \]
Multiple Choice Questions

On the ParScore form you need to fill in your answers, perm number, test version, and name. Failure to do any of these things will result in the **loss of 1 point.** Your perm number is placed and bubbled in under the “ID number.” Do not skip boxes or put in a hyphen; unused boxes should be left blank. Bubble in your test version (B) under the “test form.” Note: Your ParScore forms will not be returned to you, therefore, for your records, you may want to mark your answers on this sheet. Each multiple choice question is worth 5 points.

1. Two metals of equal mass with different heat capacities are subjected to the same amount of heat. Which undergoes the smaller change in temperature?
   A) Both undergo the same change in temperature.
   B) The metal with the higher heat capacity.
   C) To determine this, you need to know which metals you are talking about.
   D) The metal with the lower heat capacity.
   E) To determine this, you need to know the initial temperatures of the metals.

2. Calculate the $\Delta S$ associated with a process in which 5.00 mol of gas expands reversibly at constant temperature $T=25^\circ C$ from a pressure of 10.0 atm to 1.00 atm.
   A) $-95.7 \frac{J}{K}$
   B) $-28,500 \frac{J}{K}$
   C) $28,500 \frac{J}{K}$
   D) $95.7 \frac{J}{K}$
   E) None of the above

3. Choose the correct statement.
   A) At constant pressure and temperature, a decrease in free energy ensures an increase in the entropy of the system.
   B) A reaction that exhibits a negative value of $\Delta S$ cannot be spontaneous.
   C) Free energy is independent of temperature.
   D) Exothermic reactions are always spontaneous.
   E) None of the above

4. Which of the following statements is(are) true?
   A) The heat of reaction and change in enthalpy can always be used interchangeably.
   B) Enthalpy is a state function.
   C) In exothermic reactions, the reactants are lower in potential energy than the products.
   D) A chemist takes the point of view of the surroundings when determining the sign for work or heat.
   E) At least two of these statements are true.
5. In a system composed of nitrogen gas in a cylinder fitted with a piston, when 2.00 kJ of energy is transferred into the system as heat at an external pressure of 2.00 atm, the nitrogen gas expands from 2.00 L to 5.00 L against the constant pressure. What is $\Delta E$ for the process?
   A) -0.608 kJ
   B) 1.39 kJ
   C) 0 kJ
   D) 2.61 kJ
   E) None of the above

Consider a process carried out on 1.00 mol of a monatomic ideal gas by the following two different pathways. The first pathway is A (3.00 atm, 20.0 L) to C (1.00 atm, 20.0 L) to D (1.00 atm, 50.0 L); and the second pathway is A (3.00 atm, 20.0 L) to B (3.00 atm, 50.0 L) to D (1.00 atm, 50.0 L). In each case, the gas is taken from state A to state D.

6. Calculate $\Delta H_{ABD}$.
   A) 25 L$\cdot$atm
   B) 475 L$\cdot$atm
   C) -475 L$\cdot$atm
   D) -25 L$\cdot$atm
   E) None of these