Directions: (1) Put your name (neatly) and signature on both parts of the exam where indicated.  
(2) Each multiple choice question is actually 2 questions on your scanning sheet. If you are sure of an answer, put the same answer down for both questions for 5 pts. If you cannot decide between two answers, put one answer down for one question and the other answer down for the other question. If you get one correct you'll get half credit for 2.5 pts. If there is an ambiguous multiple choice question, use the last page to explain your answer. 
(3) Do NOT write on the envelope. 
(4) When finished, put both parts of the exam in the envelope with the scanning sheet. You can leave during announced times. You can pick up your multiple choice and answers outside my office after the exam. 
(5) There are a total of 29 questions (17 actual questions). 

PART 1

Please read and sign: “On my honor, as an Aggie, I have neither given nor received unauthorized aid on this exam.” 

1&2. Which of the following terms or symbols is a thermodynamic state function? 
(a) heat  
(b) work  
(c) $-\Delta n_{\text{gas}}RT$  
(d) heat + work  
(e) distance 

3&4. For which of the following substances is $\Delta G_f^\circ = 0$? 
(a) CO$_2$(g)  
(b) Ni(s)  
(c) O$_3$(g)  
(d) C (s, diamond)  
(e) I$_2$(l) 

5&6. How much heat is involved when a system does 30 J of work and experiences a gain of 20 J of internal energy? 
(a) +50 J  
(b) +20 J  
(c) +60 J  
(d) −50 J  
(e) −20 J
7&8. Consider the conversion of a substance between solid and liquid: solid $\rightarrow$ liquid. When the two phases are in equilibrium at one atmosphere pressure and at the melting point of the substance,

(a) $\Delta H = 0$ for the process.
(b) $\Delta E = 0$ for the process.
(c) $\Delta S = 0$ for the process.
(d) $\Delta G = 0$ for the process.
(e) both (a) and (b)

9&10. What is the enthalpy change of the reaction below at 298 K and 1 atm pressure?

<table>
<thead>
<tr>
<th>$\Delta H_{f}^{o}$ (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>CO$_2$(g) + 2H$_2$S(g) $\rightarrow$ CS$_2$(l) + 2H$_2$O(l)</td>
</tr>
<tr>
<td>−394.0</td>
</tr>
<tr>
<td>−20.2</td>
</tr>
<tr>
<td>+89.5</td>
</tr>
<tr>
<td>−286.0</td>
</tr>
</tbody>
</table>

(a) −472.5 kJ  (b) −68.3 kJ  (c) −896.7 kJ  (d) −264.1 kJ  (e) −48.1 kJ

11&12. Under conditions of constant volume, the heat change that occurs during a chemical reaction is equal to:

(a) $\Delta H$  (b) $\Delta E$  (c) $\Delta T$
(d) $\Delta P$  (e) $\Delta w$
13&14. When 1.00 g of magnesium is burned at 25°C and constant pressure to give MgO, 24.8 kJ of heat energy are liberated.

\[ \text{Mg(s)} \text{ + 1/2 O}_2(g) \rightarrow \text{MgO(s)} \]

The standard enthalpy of formation of MgO is:

(a) −24.8 kJ/mol  (b) +24.8 kJ/mol  (c) −999 kJ/mol  (d) +999 kJ/mol  (e) −602 kJ/mol

15&16. Which of the following statements is FALSE?

(a) Endothermic processes are those with $\Delta H > 0$.
(b) $\Delta H$ is a state function.
(c) Some spontaneous processes are endothermic.
(d) Entropy alone determines the spontaneity of a reaction.
(e) $\Delta G$ represents the free energy change for a process.
17&18. Consider the following reaction:

\[ 2 \text{Na(s)} + 2 \text{H}_2\text{O(l)} \rightarrow \text{H}_2(g) + 2 \text{NaOH(aq)} \quad \Delta H = -368 \text{kJ/mol rxn} \]

How much heat is evolved if 32.0 g of sodium metal react with excess water?

(a) 8.0 kJ               (b) 124 kJ               (c) 256 kJ
(d) 1040 kJ             (e) 13500 kJ

19&20. For the following reaction at 25°C, \( \Delta H^\circ = -35.9 \text{kJ} \) and \( \Delta S^\circ = +446 \text{J/K} \).

Calculate \( \Delta G^\circ \) for the reaction at 25°C in kJ.

\[ \text{NH}_4\text{NO}_3(s) \rightarrow \text{N}_2\text{O(g)} + 2 \text{H}_2\text{O(g)} \]

(a) \(-169 \text{kJ}\)               (b) +97. kJ               (c) \(-24.8 \text{kJ}\)
(d) +10.7 kJ            (e) +47.1 kJ
21&22. Determine the enthalpy change for the formation of methane from solid carbon (as graphite) and hydrogen gas,

\[ \text{C(s)} + 2 \text{H}_2(\text{g}) \rightarrow \text{CH}_4(\text{g}) \]

From the following equations:

<table>
<thead>
<tr>
<th>Equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{CO}_2(\text{g}) \rightarrow \text{C(s)} + \text{O}_2(\text{g}) )</td>
</tr>
<tr>
<td>( \Delta H^\circ = +393.5 \text{ kJ} )</td>
</tr>
<tr>
<td>( 4 \text{H}_2(\text{g}) + 2 \text{O}_2(\text{g}) \rightarrow 4 \text{H}_2\text{O}(\ell) )</td>
</tr>
<tr>
<td>( \Delta H^\circ = -1143.2 \text{ kJ} )</td>
</tr>
<tr>
<td>( \text{CH}_4(\text{g}) + 2 \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\ell) )</td>
</tr>
<tr>
<td>( \Delta H^\circ = -890.3 \text{ kJ} )</td>
</tr>
</tbody>
</table>

(a) \(-635 \text{ kJ}\)  (b) \(+204 \text{ kJ}\)  (c) \(-74.8 \text{ kJ}\)  (d) \(-1640 \text{ kJ}\)  (e) \(+74.8 \text{ kJ}\)

23&24. Consider the following reaction and choose the correct statement:

\[ \text{H}_2(\text{g}) + \text{Br}_2(\ell) \rightarrow 2 \text{HBr}(\text{g}) \]

\( \Delta H^\circ = -72.8 \text{ kJ} \)

\( \Delta S^\circ = +114 \text{ J/K} \)

(a) The reaction is nonspontaneous at all temperatures.
(b) The reaction is spontaneous at all temperatures.
(c) The reaction becomes spontaneous as the temperature increases.
(d) The reaction becomes nonspontaneous as the temperature increases.
(e) The temperature does not influence the spontaneity of any reaction.
25. Consider the reaction: \(2H_2(g) + O_2(g) \rightarrow 2H_2O(l)\)

(a) Draw a particle view of the reaction, using \(\bigcirc\) for a hydrogen atom and \(\bigcirc\) for an oxygen atom.

(b) Calculate the heat generated when 0.0200 moles of hydrogen gas reacts with 0.0100 moles of oxygen gas in a bomb calorimeter containing 544 grams of water. The heat capacity of the calorimeter is 125 J/oC. The temperature of the water in the calorimeter rose from 24.000°C to 26.381°C. The specific heat of water is 4.184 J/g·oC.

(c) What is the \(\Delta E_{\text{rxn}}\) (kJ/mol O₂)?

(d) What is the \(\Delta H_{\text{rxn}}\) (kJ/mol O₂)? Use the final T.
(5 pts) (e) Determine the $\Delta H^\circ$ of formation of H$_2$O($l$), using your results from (d), and explain your answer. If you didn't get that far, use $\Delta H^\circ = -100$ kJ/mol of O$_2$ as your answer to (d).

(5 pts) (f) Looking at the reaction and the particle view, briefly discuss the entropy of this reaction. What sign should it have? Include the sign convention for entropy and the usual units.

(5 pts) (h) This reaction was spontaneous at the temperature of this experiment because it occurred. If you raised the temperature, could the reaction become nonspontaneous? Explain your answer. (Remember, you've calculated $\Delta H$ of the reaction and you've determined the sign of $\Delta S$.) Explain using the Gibbs-Helmholtz equation and sign conventions to get full credit.
(5 pts) 26. Consider the work involved when 1 mol of gaseous ethane is completely combusted at 25°C (Hint - is water a liquid or solid at this temperature?).

(i) Write the balanced equation.
   If you don't know the formula for ethane, use C₆H₁₄(ℓ) and lose 3 pts.

(ii) Circle the correct answer below. To get any credit, I must see your balanced equation.

(a) $$\Delta n_{\text{gas}}$$ is negative; so work is done on the system by the surroundings.
(b) $$\Delta n_{\text{gas}}$$ is negative; so work is done by the system on the surroundings.
(c) $$\Delta n_{\text{gas}}$$ is positive; so work is done on the system by the surroundings.
(d) $$\Delta n_{\text{gas}}$$ is positive; so work is done by the system on the surroundings.
(e) $$\Delta n_{\text{gas}}$$ is zero, so no work is done.

27. Consider the following data for pure liquid nitric acid.

Pure nitric acid, HNO₃, boils at 83°C:  $$\text{HNO}_3(\ell) \rightleftharpoons \text{HNO}_3(g)$$

Given: $$\Delta H_{298}^\circ$$ (kJ/mol) for HNO₃(ℓ) = −174.1 kJ/mol and $$\Delta H_{298}^\circ$$ (kJ/mol) for HNO₃(g) = −135 kJ/mol.

(3 pts) (a) Calculate the change in entropy, $$\Delta S$$, for this process. Give the proper units.

(2 pts) (b) Explain why the sign of your answer makes sense.