Statistics:
- **Average:** 43 pts (62%);
- **Highest:** 67 pts (96%); **Lowest:** 10 pts (14%)
- Number of students performing at or above average: 35 (51%)

**PART I: THERMOCHEMISTRY**

1. (5 pts) Mark as true (T) or false (F) the following statements. Do not explain!
   - (T) Enthalpy is a state function;
   - (F) Reactions resulting in the release of heat are called endothermic;
   - (F) The heat transfer at constant pressure is identical with the energy change of the system;
   - (F) Fuel value is defined as the energy released from the reaction of exactly 1 mole of fuel;
   - (T) Heat of vaporization and heat of condensation have the same absolute values but opposite signs;

2. When 1.0 g of certain substance is heated, the following graph results (S = solid; L = liquid; V = vapor):

   ![Graph]

   Answer the following questions:
   A. (2 pts) Which quantity has larger absolute value: Heat of condensation or heat of freezing?
   
   **Heat of condensation**

   B. (3 pts) Calculate the heat change during condensation of 2.5 g of V at 140 °C;

   \[ 2.5 \text{ g} \times (-200 \text{ J/g}) = -500 \text{ J} \]

   C. (2 pts) Which has larger specific heat capacity: S or L?

   L

   D. (4 pts) What quantity of heat is required to raise the temperature of 3.8 g of L by 6 °C?

   **Answer:** From the plot of L we can derive the specific heat capacity: 200 J/(60°C x 1.0 g) = 3.33 J/g°C

   \[ q = m \times c \times \Delta T = 3.8 \text{ g} \times 3.33 \text{ J/g°C} \times 6^\circ\text{C} = 76 \text{ J} \]

3. A quantity of 0.548 g of the sugar arabinose, C₅H₁₀O₅, is burned completely in oxygen, in a calorimeter to give carbon dioxide and water. The calorimeter’s temperature rose from 20.00 °C to 20.54 °C. The heat capacity of the calorimeter is 15.8 J/°C.

   A. (5 pts) Calculate the amount of heat transfer in the above reaction. Is the reaction exothermic or endothermic?

   **Answer:** Since the heat capacity (NOT specific heat capacity!) of the calorimeter is given, we can calculate directly the amount of heat:

   \[ q = \text{heat capacity} \times \Delta T = 15.8 \text{ J/°C} \times 0.54 \text{ °C} = 8.5 \text{ J} \]
\[ q_{\text{reaction}} = - q_{\text{calorimeter}} = -8.5 \text{ J} \] - the reaction is exothermic

B. (4 pts) Write the thermochemical equation for the burning of 1 mole of arabinose;

**Answer:** The above calculated amount of heat is for 0.548 g. We need to find the amount per mole of arabinose. The molar mass is approximately 150. So the following relationship holds:

\[-8.5 \text{ J}/0.548 \text{ g} = \Delta H^o/150 \text{ g}\]

\[\Delta H^o = -2.3 \text{ kJ}\]

C. (3 pts) Calculate the fuel value of arabinose;

**Answer:** The fuel value is defined as the amount of heat liberated per gram of fuel. Therefore:

\[2300 \text{ J}/150 \text{ g} = \Delta H/1 \text{ g}\]

\[\Delta H = \text{fuel value} = 15.3 \text{ J/g}\]

4. (6 pts) Hydrogen cyanide (HCN) is a highly poisonous, volatile liquid. It can be prepared according to the reaction:

\[\text{CH}_4(g) + \text{NH}_3(g) \rightarrow \text{HCN}(g) + 3 \text{H}_2(g)\]

Calculate the heat of this reaction at constant pressure, using the following information:

\[\text{N}_2(g) + 3 \text{H}_2(g) \rightarrow \text{NH}_3(g) \quad \Delta H = -91.8 \text{ kJ}\]

\[\text{C(graphite)} + 2 \text{H}_2(g) \rightarrow \text{H}_2(g) \quad \Delta H = -74.9 \text{ kJ}\]

\[\text{H}_2(g) + 2 \text{C(graphite)} + \text{N}_2(g) \rightarrow \text{HCN}(g) \quad \Delta H = 270.3 \text{ kJ}\]

**Answer:** The above equations can be manipulated to give us, when added, the equation in question: The first equation needs to be reversed and divided by 2; the second has to be reversed; the third has to be divided by 2:

\[\text{NH}_3(g) \rightarrow ½ \text{N}_2(g) + 3/2 \text{H}_2(g) \quad \Delta H = 45.9 \text{ kJ}\]

\[\text{CH}_4(g) \rightarrow \text{C(graphite)} + 2 \text{H}_2(g) \quad \Delta H = 74.9 \text{ kJ}\]

\[½ \text{H}_2(g) + \text{C(graphite)} + ½ \text{N}_2(g) \rightarrow \text{HCN}(g) \quad \Delta H = 135.2 \text{ kJ}\]

\[\text{NH}_3(g) + \text{CH}_4(g) + ½ \text{H}_2(g) + \text{C(graphite)} + ½ \text{N}_2(g) \rightarrow ½ \text{N}_2(g) + 3/2 \text{H}_2(g) + \text{C(graphite)} + 2 \text{H}_2(g) + \text{HCN}(g)\]

\[\text{CH}_4(g) + \text{NH}_3(g) \rightarrow \text{HCN}(g) + 3 \text{H}_2(g) \quad \Delta H = 45.9 \text{ kJ} + 74.9 \text{ kJ} + 135.2 \text{ kJ} = 256.0 \text{ kJ}\]

5. (6 pts) Ammonia burns in oxygen according to the equation:

\[4 \text{NH}_3(g) + 5 \text{O}_2(g) \rightarrow 4 \text{NO}(g) + 6 \text{H}_2\text{O}(g) \quad \Delta H^o = -906 \text{ kJ}\]

How much heat is released when a mixture containing 10.0 g of NH\textsubscript{3} and 20.0 g of O\textsubscript{2} is reacted? (**Hint:** You have to decide on a limiting reactant first!!!)

**Answer:** According to the stoichiometry of the above equation, we have 10.0 g x 1/17.0 g/mol = 0.588 mol of NH\textsubscript{3}, and it reacts completely with:

0.588 mol x (5 mol O\textsubscript{2}/4 mol N\textsubscript{2}) x 32 g/mol = 23.5 g of O\textsubscript{2}

But there are only 20.0 g of O\textsubscript{2}, which means that O\textsubscript{2} is the limiting reagent
According to the thermochemical equation:

\[- 906 \text{ kJ}/5 \text{ mol } \text{O}_2 = - 906 \text{ kJ}/160 \text{ g } \text{O}_2 = \Delta H/20.0 \text{ g } \text{O}_2\]

\[\Delta H = - 113 \text{ kJ}\]

6. Carbon disulfide, a colorless liquid, burns in oxygen according to the unbalanced equation:

\[
\text{CS}_2(\text{l}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{SO}_2(\text{g})
\]

A. (5 pts) Write a balanced thermochemical equation for the reaction above (Standard molar enthalpies of formation: CS\(_2\)(l): 87.9 kJ/mol; CO\(_2\)(g): - 393.5 kJ/mol; SO\(_2\)(g): - 296.8 kJ/mol).

**Answer:** First we have to balance the equation:

\[
\text{CS}_2(\text{l}) + 3 \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{SO}_2(\text{g})
\]

According to the general formula, relating heats of reactions to heats of formation:

\[
\Delta H = \sum (\text{# mole x } \Delta H_{\text{f}})_{\text{products}} - \sum (\text{# mol x } \Delta H_{\text{f}})_{\text{reactants}} = - 393.5 + 2 \times (-296.8) - 87.9 = - 1075 \text{ kJ}
\]

B. (4 pts) What is the enthalpy change when 120 kg of CS\(_2\)(l) is burned?

\[120 \text{ kg of } \text{CS}_2 = 120000 \text{ g} = 120000 \text{ g} \times \left(1 \text{ mol}/76.1 \text{ g}\right) = 1577 \text{ mol of } \text{CS}_2\]

Therefore: \[- 1075 \text{ kJ}/1 \text{ mol} = \Delta H/1577 \text{ mol}\]

\[\Delta H = - 1695 \text{ megaJ}\]

PART II: ELECTRONIC STRUCTURE.

**Relevant constants:**  
Plank constant: 6.626 \times 10^{-34} \text{ J.s}  
Rydberg constant: 2.179 \times 10^{-18} \text{ J}  
Speed of light: 3 \times 10^8 \text{ m/s}  
Mass of electron: 9.11 \times 10^{-31} \text{ kg}

7. The green line in the line spectrum of the element thallium has a wavelength of 535 nm.

A. (4 pts) Calculate the energy of a photon of this light.

**Answer:**  
\[E = h\nu = \frac{hc}{\lambda} = \left(6.626 \times 10^{-34} \text{ J.s} \times 3 \times 10^8 \text{ m/s}\right)/535 \times 10^{-9} \text{ m} = 3.72 \times 10^{-19} \text{ J}\]

B. (3 pts) Calculate the energy of 0.25 mol of photons of this light.

**Answer:**  
The value calculated in (A) is for 1 photon. Therefore, for 0.25 mol of photons:  
\[0.25 \times 6.022 \times 10^{23} \times 3.72 \times 10^{-19} \text{ J} = 5.6 \times 10^4 \text{ J} = 56 \text{ kJ}\]

8. (6 pts) Calculate the shortest wavelength of the electromagnetic radiation emitted by a hydrogen atom undergoing a transition from the \(n = 6\) level.

**Answer:**  
We need to use Bohr’s formula for energy transitions. The shortest wavelength would (according to Plank’s formula!) correspond to the highest energy transition. From any level, this would be the transition to the level \(n = 1\). Therefore:
\[ \Delta E = - R_\text{th}(1/n_1^2 - 1/n_2^2) = 2.179 \times 10^{-18} \text{ J} \]  
(1/36 - 1/1) = - 2.12 \times 10^{-18} \text{ J} 

The value is negative, indicating that the system’s energy has decreased (Light has been emitted!).

The energy of the emitted photon will have the same magnitude, but opposite sign.

\[ E_{\text{photon}} = 2.12 \times 10^{-18} = \frac{hc}{\lambda} \]

\[ \lambda = \frac{hc}{E_{\text{photon}}} = \frac{(6.62 \times 10^{-34} \text{ J s} \times 3 \times 10^8 \text{ m/s})}{2.12 \times 10^{-18} \text{ J}} = 9.38 \times 10^{-8} \text{ m} = 94 \text{ nm} \]

9. (8 pts) Each of the following sets of quantum numbers is impossible. Indicate the wrong value and suggest a correct one. (NOTE: There may be more than one correct value! Please write only one!). Give the notation for the subshell (using a correct letter designation for l) in each case.

A. 1, 0, 0, +1

\( m_s \) has only two allowed values: +1/2 or -1/2

An allowed combination would be 1, 0, 0, +1/2

Subshell: 1s

B. 3, 2, 3, -1/2

For \( l = 2 \), \( m_l \) cannot be 3. Allowed values are -2, -1, 0, 1, 2

An allowed combination would be 3, 2, 0, -1/2

Subshell: 3d

C. 0, 1, 0, +1/2

\( n \) cannot be 0. The lowest possible value is 1, but in this case it has to be at least 2

An allowed combination would be 2, 1, 0, +1/2

Subshell: 2p

D. 2, 2, -1, -1/2

For \( n = 2 \), \( l \) cannot be 2. Allowed values are 0 and 1

An allowed combination would be 2, 1, -1, -1/2

Subshell: 2p

10. (5 pts) BONUS PROBLEM!! (You do not have to solve this one, in order to get full credit). Acetone, a common solvent, burns according to the equation:

\[ \text{C}_3\text{H}_6\text{O(l)} + 4 \text{O}_2\text{(g)} \rightarrow 3 \text{CO}_2\text{(g)} + 3 \text{H}_2\text{O(l)} \quad \Delta H^\circ = -1791 \text{ kJ} \]

Calculate the standard molar enthalpy of formation of acetone, \( \text{C}_3\text{H}_6\text{O(l)} \) (Other values that you will need: \( \Delta H^\circ (\text{CO}_2\text{(g)}) = -393.5 \text{ kJ/mol} \);
\( \Delta H^\circ (\text{H}_2\text{O(l)}) = -285.8 \text{ kJ/mol} \)

**Answer:** According to the general formula, relating heats of reactions to heats of formation:

\[ \Delta H^\circ = \Sigma ( \# \text{ mole} \times \Delta H^\circ)_{\text{products}} - \Sigma ( \# \text{ mol} \times \Delta H^\circ)_{\text{reactants}}\]

There are two reactants, but oxygen is an element and its heat of formation is 0 by definition. Therefore:

\[ -1791 \text{ kJ} = 3 \times (-285.8 \text{ kJ}) + 3 \times (-393.5 \text{ kJ}) - \Delta H^\circ (\text{acetone}) \]

\[ - \Delta H^\circ (\text{acetone}) = 246.9 \text{ kJ} \]

\[ \Delta H^\circ (\text{acetone}) = - 246.9 \text{ kJ} \]