I. Terms, short answers, and graphs. (Total 5 pts, 1 pt each)

1. The amount of heat required to heat one gram of a substance by one degree is called

   **Specific heat capacity**

2. The statement of the first law of thermodynamics is

   **The internal energy of an isolated system does not change \( \Delta U=0 \)**

3. Heat of combustion is measured with the instrument called

   **Bomb calorimeter**

4. The enthalpy for the reaction \( \text{F(g)} + \text{e}^- (g) \rightarrow \text{F}^-(g) \) is called

   **Electron gain enthalpy**

5. Sketch the enthalpy change diagram for the following reaction:

   \[
   \text{C}_2\text{H}_6(g) + \frac{7}{2}\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 3\text{H}_2\text{O}(l) \quad \Delta H = -1560 \text{ kJ/mol}
   \]

   ![Enthalpy change diagram for the reaction \( \text{C}_2\text{H}_6(g) + \frac{7}{2}\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 3\text{H}_2\text{O}(l) \)]
II. Calculations of $\Delta H$, $\Delta U$, q, and w. (8 pts, 4 pts each)

Calculate the quantities $\Delta H$, $\Delta U$, q, and w for the following processes. Be sure to give signs! Gases may be considered ideal, and the volumes of liquids and solids (but not gases) may be ignored.

<table>
<thead>
<tr>
<th>initial state</th>
<th>final state</th>
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<tbody>
<tr>
<td>(1) 2NH$_3$ (g), 300 K, 1 bar</td>
<td>2NH$_3$ (g), 300 K, 5 bar</td>
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with the external pressure 10 bar

$\Delta U = \Delta H = 0$ (For an isothermal compression of an ideal gas)

$\Delta U = q + w$  \Rightarrow  $q = -w$ ($\Delta U = 0$)

$w = -p_{\text{ext}} \Delta V$ (the sign is chosen so that work is done on the system (positive))

$\Delta V$ is calculated using $pV = nRT$:

\[
\begin{align*}
n &= 2 \text{ mol} \\
R &= 8.3145 \text{ J K}^{-1} \text{ mol}^{-1} \\
T &= 300 \text{ K} \\
p_{\text{init}} &= 1 \times 10^5 \text{ Pa} \\
p_{\text{fin}} &= 5 \times 10^5 \text{ Pa}
\end{align*}
\]

$\Delta V = -0.0399 \text{ m}^3$  $p_{\text{ext}} = 10 \times 10^5 \text{ Pa}$  \Rightarrow  $w = 39 \text{ kJ}$

$q = -w = -39 \text{ kJ}$

(2) CH$_3$OH(l) 350K, 1 bar  \rightarrow  CH$_3$OH(s) 150K, 1 bar

$C_p[\text{CH}_3\text{OH}(s)] = 110 \text{ J/(K mol)}$  $C_p[\text{CH}_3\text{OH}(l)] = 81.6 \text{ J/(K mol)}$

$\Delta_{\text{fus}}H^0(\text{CH}_3\text{OH}) = 3.16 \text{ kJ/mol}$  $T_{\text{fus}} = 175 \text{ K}$

We are dealing with solids and liquids here. Therefore, the volume changes are negligible and the work is zero:

$\Delta V = 0$  \quad $w = 0$

Also, the pressure changes can be ignored, so $\Delta U = \Delta H$: 
\[ \Delta U = \Delta H - \Delta(pV) \quad \Delta(pV) = 0 \]

We calculate the heat for three steps: 1) Cooling from the initial temperature to the fusion temperature. 2) Forming the solid from the liquid. 3) Cooling the solid to the final temperature.

\[ q_1 = nC\Delta T = (1 \text{ mol})(81.6 \text{ J K}^{-1} \text{ mol}^{-1})(-175 \text{ K}) = -14.28 \text{ kJ} \]
\[ q_2 = n\Delta H_{\text{fus}} = (1 \text{ mol})(3.16 \text{ kJ/mol}) = 3.16 \text{ kJ} \]
\[ q_3 = nC\Delta T = (1 \text{ mol})(110 \text{ J K}^{-1} \text{ mol}^{-1})(-25 \text{ K}) = -2.75 \text{ kJ} \]
\[ q_{\text{tot}} = (-14.28 - 3.16 - 2.75) \text{ kJ} = -20.2 \text{ kJ} \]
\[ \Delta U = \Delta H = q = -20.2 \text{ kJ} \]

III. Problems  [total 7 pts, (1) 3 pts. (2) 4 pts]

(1) The standard enthalpy of formation of propene \( \text{C}_3\text{H}_6(g) \) is +20.42 kJ/mol and the standard enthalpy of formation of cyclopropane is +53.30 kJ/mol. Using this information calculate the standard enthalpy of isomerization of cyclopropane.

\[ \text{Cyclopropane} \rightarrow \text{Propene} \]
\[ \Delta H_{\text{rxn}} = \Delta H_{f, \text{propene}} - \Delta H_{f, \text{cyclopropane}} \]
\[ = (20.42 \text{ kJ/mol}) - (53.30 \text{ kJ/mol}) \]
\[ = -32.88 \text{ kJ/mol} \]

Reversing the reaction will change the sign of the final answer.

(2) When 0.3212 g of glucose, \( \text{C}_6\text{H}_{12}\text{O}_6 \), was burned in a bomb calorimeter with the calorimeter constant 641 J/K the temperature rose by 7.793 K. Calculate standard molar internal energy of combustion and standard molar enthalpy of combustion of glucose.

\[ M_w(\text{C}_6\text{H}_{12}\text{O}_6) = 180.16 \text{ g/mol} \]

The rxn of interest is:
\[
\text{C}_6\text{H}_{12}\text{O}_6\text{(s)} + 6 \text{O}_2\text{(g)} \rightarrow 6 \text{CO}_2\text{(g)} + 6 \text{H}_2\text{O}\text{(l)}
\]

0.3212 g of sucrose = 0.00178 mol of sucrose

The heat given off by the combustion is:

\[q = -(641 \text{ J/K})(7.793 \text{ K}) = -5.00 \text{ kJ} \text{ or } -2810 \text{ kJ/mol}\]

Since this is a constant volume process, \(\Delta U = q\). We also make the approximation that ‘T’ is constant for the sake of calculating \(\Delta H\). Though not entirely correct, it does not introduce a large error and we justify it by saying that we are interested in \(\Delta H\) at a particular temperature. Points were not deducted if you didn’t assume this. Using this assumption:

\[\Delta(pV) = \Delta(nRT) = nR\Delta T = 0\]

(Note that \(\Delta n = 0\) because the amount of gases present does not change. We ignore solids and liquids.)

Therefore, we have \(\Delta H = \Delta U\):

\[\Delta H = \Delta U = q = -2810 \text{ kJ/mol}\]