Using the following data:

<table>
<thead>
<tr>
<th></th>
<th>ΔH° (kJ/mol)</th>
<th>S° (J K⁻¹ mol⁻¹)</th>
</tr>
</thead>
<tbody>
<tr>
<td>N₂O₅ (g)</td>
<td>11.29</td>
<td>355.3</td>
</tr>
<tr>
<td>NO₂ (g)</td>
<td>33.15</td>
<td>239.9</td>
</tr>
<tr>
<td>O₂ (g)</td>
<td>0</td>
<td>204.8</td>
</tr>
</tbody>
</table>

1. (4 pts) Calculate ΔH° for the reaction 2 N₂O₅ (g) → 4 NO₂ (g) + O₂ (g) Circle the answer.
   
   a) −110.02 kJ
   b) +110.02 kJ
   c) −155.20 kJ
   d) +155.20 kJ
   e) −21.86 kJ
   f) +21.86 kJ
   
   ![Image]

   \[ \Delta H^o = 4(33.15) - 2(11.29) \]
   
   \[ \Delta H^o = 110.02 \text{ kJ} \]

2. (4 pts) Calculate ΔS° for the reaction 2 N₂O₅ (g) → 4 NO₂ (g) + O₂ (g) Circle the answer.
   
   a) 89.5 J/K
   b) 249.2 J/K
   c) 453.8 J/K
   d) 249.2 J/K
   e) −115.6 J/K
   
   ![Image]

   \[ \Delta S^o = 4(239.9) + 204.8 - 2(355.3) \]
   
   \[ \Delta S^o = 453.8 \text{ J/K} \]

3. (6 pts) In class we observed the following reaction at room temperature and pressure:

   \[ \text{Ba(OH)}_2 + 8 \text{H}_2\text{O} (s) + 2 \text{NH}_4\text{SCN} (s) \rightarrow \text{Ba(SCN)}_2 (s) + 2 \text{NH}_3 (g) + 10 \text{H}_2\text{O} (l) \]

   The solution produced was so cold that it froze the water underneath the beaker, adhering a block of wood to the beaker. Why is this reaction spontaneous? Circle the correct answer.
   
   a) The reaction is spontaneous because the reaction is exothermic.
   b) The reaction is spontaneous because ΔH > 0.
   c) The reaction is spontaneous because the entropy increases.
   d) The reaction is spontaneous because ΔS < 0.

4. (6 pts) Given the following data at 25°C:

   \[ \text{CH}_4 (g) + 2 \text{O}_2 (g) \rightarrow 2 \text{CO}_2 (g) + \text{H}_2\text{O} (g) \quad \Delta H^o = -981.1 \text{ kJ} \]
   
   \[ \text{CH}_4 (g) + 2 \text{O}_2 (g) \rightarrow \text{CO}_2 (g) + 2 \text{H}_2\text{O} (g) \quad \Delta H^o = -802.3 \text{ kJ} \]

   Calculate ΔH° for the following reaction at 25°C:

   \[ 2 \text{CH}_4 (g) + 2 \text{O}_2 (g) \rightarrow \text{CH}_2\text{CO} (g) + 3 \text{H}_2\text{O} (g) \]

   \[ 2 \text{CO}_2 + \text{H}_2\text{O} \rightarrow \text{CH}_2\text{CO} + 2 \text{O}_2 \quad +981.1 \text{ kJ} \]
   
   \[ 2 [\text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O}] \quad 2 (-802.3 \text{ kJ}) \]

   \[ 2 \text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CH}_2\text{CO} + 3 \text{H}_2\text{O} \quad -623.5 \text{ kJ} \]
5. (6 pts) For the reaction \( A + B \rightarrow C + D, \quad \Delta H^\circ = +40 \text{ kJ} \) and \( \Delta S^\circ = 50 \text{ J/K}. \)

Therefore, under standard conditions, the reaction

- \( a) \) is spontaneous at all temperatures.
- \( b) \) is spontaneous at temperatures less than 10 K.
- \( c) \) is spontaneous at temperatures greater than 800 K.
- \( d) \) is spontaneous at temperatures less than 800 K.
- \( e) \) is not spontaneous at any temperature.

\[ \Delta G = \Delta H - T \Delta S \]
\[ 0 = 40 \text{ kJ} - T(50 \text{ J/K})(1000 \text{ J}) \]
\[ T = 800 \text{ K} \]

At \( T > 800 \text{ K} \) \( \Rightarrow \Delta G < 0 \)

6. (6 pts) At constant temperature and pressure, the dissociation of hydrogen, \( \text{H}_2(\text{g}) \rightarrow 2 \text{H}(\text{g}) \)

- \( a) \) is spontaneous at any temperature.
- \( b) \) is spontaneous at high temperature.
- \( c) \) is spontaneous at low temperatures.
- \( d) \) is not spontaneous at any temperature.

\[ \Delta G = \Delta H - T \Delta S \]
\[ \Delta G < 0 \]

\[ \Delta H > 0 \]

Breaking \( \text{H-H} \) bond requires energy \( \Rightarrow \Delta H > 0 \)

7. (8 pts) Using the following data:

<table>
<thead>
<tr>
<th>Substance</th>
<th>( \Delta H^\circ \text{ (kJ/mol)} )</th>
</tr>
</thead>
<tbody>
<tr>
<td>CO (g)</td>
<td>-110.5</td>
</tr>
<tr>
<td>Pb (s)</td>
<td>0</td>
</tr>
<tr>
<td>CO(_2) (g)</td>
<td>-393.5</td>
</tr>
</tbody>
</table>

Calculate \( \Delta H^\circ \) for the PbO (s).

\[ \text{PbO (s) + CO (g) \rightarrow Pb (s) + CO}_2 \text{(g)} \]

\[ \Delta H^\circ = -131.4 \text{ kJ} \]

\[ -393.5 = [\Delta H^\circ (\text{PbO (s)}) + (-110.5)] \]

\[ \Delta H^\circ = -151.6 \text{ kJ/mol} \]

8. (8 pts) The standard enthalpy of formation of H\(_2\)O (l) at 298 K is \(-285.6 \text{ kJ/mol}\). Calculate \( \Delta E \) for the following process at 298 K.

\[ \text{H}_2(\text{g}) + (1/2)\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O (l)} \quad \Delta H = -285.6 \text{ kJ} \]

- \( a) \) \(-278.2 \text{ kJ/mol}\)
- \( b) \) \(-281.9 \text{ kJ/mol}\)
- \( c) \) \(-285.6 \text{ kJ/mol}\)
- \( d) \) \(-289.3 \text{ kJ/mol}\)
- \( e) \) \(3430 \text{ kJ/mol}\)

\[ \Delta E = q + w \]

\[ \Delta E = \Delta H - \Delta n_{\text{gas}} RT \]

\[ \Delta E = -285.6 \text{ kJ} - (-3/2 \text{ mol})(8.3145 \text{ J/mol K})(298 \text{ K})(1000 \text{ J}) \]

\[ \Delta E = -281.9 \text{ kJ} \]
9. (6 pts) Consider the reaction, \( A \rightleftharpoons B \) at equilibrium. Given the following relationship:

\[ \ln K = -\frac{\Delta H^\circ}{RT} + \frac{\Delta S^\circ}{R} \]

If the reaction is exothermic, what is the effect on the equilibrium constant when the temperature is increased? Assume \( \Delta H^\circ \) and \( \Delta S^\circ \) are constant over the change in temperature. Circle the correct answer.

a) No effect on the equilibrium constant \( K \)

b) \( K \) decreases

c) \( K \) increases

10. (6 pts) Calculate the enthalpy change when 2.38 g of CO vaporizes at its normal boiling point. Given that \( \Delta H_{vap} = 6.04 \, kJ / \text{mol} \).

\[ \text{CO (l)} \rightarrow \text{CO (g)} \quad \Delta H_{vap} = 6.04 \, \text{kJ/mol} \]

\[ \frac{2.38 \, \text{g}}{28 \, \text{g/mol}} = 0.085 \, \text{mol CO} \]

\[ \Delta H = \frac{6.04 \, \text{kJ/mol}}{1 \, \text{mol CO}} \]

\[ \Delta H = 0.51 \, \text{kJ} \]

11. (4 pts) Choose the correct statement.

a) Exothermic reactions are always spontaneous.

b) Free energy is independent of temperature.

c) A reaction that exhibits a negative value of \( \Delta S \) can not be spontaneous.

d) At constant pressure and temperature, a decrease in free energy ensures an increase in the entropy of the system.

e) None of these statements are correct.

12. (4 pts) For which process is \( \Delta S \) negative?

a) evaporation of 1 mol of CCl\(_4\) (l)

b) mixing 5 mL ethanol with 25 mL water

c) compressing 1 mol of Ne at constant temperature from 0.5 atm to 1.5 atm.

d) grinding a large crystal of KCl to powder

e) raising the temperature of 100 g of Cu (s) from 275 K to 295 K

13. (6 pts) For the freezing of liquid water at \(-10^\circ\text{C}\) and 1 atm pressure, circle the correct answer for \( \Delta H \), \( \Delta S \) and \( \Delta G \).

\[ \Delta H < 0 \quad \Delta H > 0 \quad \Delta H = 0 \]

\[ \Delta S < 0 \quad \Delta S > 0 \quad \Delta S = 0 \]

\[ \Delta G < 0 \quad \Delta G > 0 \quad \Delta G = 0 \]

Hint: First write the reaction and think about the process. Is heat consumed or produced? Does the entropy increase or decrease? Is the process spontaneous at \(-10^\circ\text{C}\) and 1 atm?

\[ \text{H}_2\text{O (l)} \rightarrow \text{H}_2\text{O (s)} \quad -10^\circ\text{C}, \quad 1 \text{ atm} \]

\( \Delta G = \Delta H - T \Delta S \)

\( \Delta G < 0 \) under these conditions.
14. (12 pts) A gas expands isothermally and irreversibly against a constant non-zero pressure.

Circle the correct answer for $w$, $q$, $\Delta H$, $\Delta E$, $\Delta S$ and $\Delta G$.

$$\Delta T = 0 \Rightarrow \Delta E = 0, \Delta H = 0$$

- $w < 0$
- $w > 0$
- $w = 0$

- $q < 0$
- $q > 0$
- $q = 0$

- $\Delta H < 0$
- $\Delta H > 0$
- $\Delta H = 0$

- $\Delta S < 0$
- $\Delta S > 0$
- $\Delta S = 0$

- $\Delta G < 0$
- $\Delta G > 0$
- $\Delta G = 0$

It is very important to write down the equations that you need to determine the correct sign of each of the quantities.

$$\Delta v > 0 \Rightarrow \Delta S > 0$$

$$w = -P_{ext} \Delta v \Rightarrow \Delta v > 0 \Rightarrow w < 0$$

$$q = -w \Rightarrow q > 0$$

$$\Delta G = q + w = 0$$

$$\Delta E = q + w = 0$$

15. (4 pts) At $25^\circ C$ and 1 atm, $\Delta G^o < 0$ for the reaction $2 \text{H}_2\text{S}(g) + 2 \text{SO}_2(g) \rightleftharpoons 3 \text{S}(s) + 2 \text{H}_2\text{O}(g)$

Circle the correct sign for $\Delta S^o$ and $\Delta H^o$.

- $\Delta S^o < 0$
- $\Delta S^o > 0$
- $\Delta S^o = 0$

- $\Delta H^o < 0$
- $\Delta H^o > 0$
- $\Delta H^o = 0$

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

4 mol gas

2 mol gas

solid

16. (10 pts) At $25^\circ C$, $K_p = 98.4$ for the reaction $2 \text{NO}(g) + \text{Br}_2(g) \rightleftharpoons 2 \text{NOBr}(g)$

Calculate $\Delta G$ at $25^\circ C$ when $P_{\text{NO}} = 0.2$ atm, $P_{\text{Br}_2} = 0.5$ atm, $P_{\text{NOBr}} = 0.3$ atm.

$$\Delta G = \Delta G^o + RT \ln Q$$

$$\Delta G^o = -RT \ln K$$

$$\Delta G^o = -(8.3145 \frac{\text{J}}{\text{mol K}})(298 \text{K}) \ln 98.4$$

$$\Delta G^o = -11.37 \text{kJ/mol}$$

$$\Delta G = -11.37 \frac{\text{kJ}}{\text{mol}} + (8.3145 \frac{\text{J}}{\text{mol K}})\left(\frac{\text{kJ}}{1000\text{J}}\right)(298 \text{K}) \ln \left(\frac{2^2}{(0.5)^2}\right)$$

$$\Delta G = -7.6 \frac{\text{kJ}}{\text{mol}}$$