

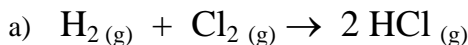
AP CHEMISTRY

TOPIC 9: THERMODYNAMICS, PART C,

Day 110:

- Bond Energy
- Third Law of Thermodynamics
- Gibbs Free Energy

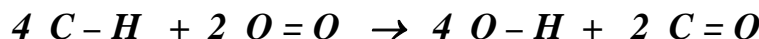
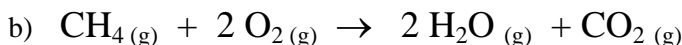
1. Using the bond energies (from the table on the examples, part II) calculate the ΔH for the reactions: You WILL NEED TO DRAW the Lewis Structures to answer these correctly.



$$\Delta H_{B.E.}^{\circ} = \sum \text{Bond energies of bonds broken (Reactants)} - \sum \text{Bond energies of bonds formed (Products)}$$

$$\Delta H_{B.E.} = \sum_{\text{Reactants}} - \sum_{\text{Products}}$$

$$\Delta H_{B.E.} = \left[\left(\frac{432 \text{ kJ}}{\text{mol}} \right) + \left(\frac{239 \text{ kJ}}{\text{mol}} \right) \right] - \left[2 \left(\frac{427 \text{ kJ}}{\text{mol}} \right) \right] = -183 \text{ kJ mol}^{-1}$$

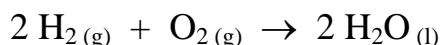


On the product side: 4 O - H, because there are 4 hydrogen atoms which form 4 single bonds to the oxygens. Also, there are two double bonds in the carbon dioxide molecule.

$$\Delta H_{B.E.} = \sum_{\text{Reactants}} - \sum_{\text{Products}}$$

$$\Delta H_{B.E.} = \left[4 \left(\frac{413 \text{ kJ}}{\text{mol}} \right) + 2 \left(\frac{495 \text{ kJ}}{\text{mol}} \right) \right] - \left[4 \left(\frac{467 \text{ kJ}}{\text{mol}} \right) + 2 \left(\frac{745 \text{ kJ}}{\text{mol}} \right) \right] = -716 \text{ kJ mol}^{-1}$$

2. Calculate the entropy change, ΔS° , for the reaction shown below.



compound	ΔS° (J mol ⁻¹ K ⁻¹)
O ₂ (g)	205.0
H ₂ O(l)	188.8
H ₂ (g)	130.6

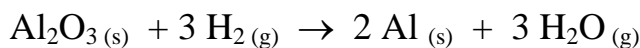
ANSWER:

$$\Delta S^{\circ} = \sum \Delta S^{\circ}(\text{products}) - \sum \Delta S^{\circ}(\text{reactants})$$

$$\Delta S^{\circ} = [2 \Delta S^{\circ}(\text{H}_2\text{O})] - [2 \Delta S^{\circ}(\text{H}_2) + \Delta S^{\circ}(\text{O}_2)]$$

$$\Delta S^{\circ} = \left[2 \left(\frac{188.8 \text{ J}}{\text{mol} \cdot \text{K}} \right) \right] - \left[2 \left(\frac{130.6 \text{ J}}{\text{mol} \cdot \text{K}} \right) + \left(\frac{205.0 \text{ J}}{\text{mol} \cdot \text{K}} \right) \right] = -88.6 \text{ J mol}^{-1} \text{K}^{-1}$$

3. Calculate the entropy change, ΔS° , for the reaction shown below.



compound	ΔS° (J mol ⁻¹ K ⁻¹)
Al ₂ O ₃ (s)	51.0
H ₂ O(g)	189.0
H ₂ (g)	130.6
Al(s)	28.0

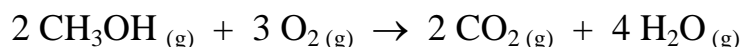
ANSWER:

$$\Delta S^\circ = \sum \Delta S^\circ (\text{products}) - \sum \Delta S^\circ (\text{reactants})$$

$$\Delta S^\circ = [2 \Delta S^\circ (\text{Al}) + 3 \Delta S^\circ (\text{H}_2\text{O})] - [\Delta S^\circ (\text{Al}_2\text{O}_3) + 3 \Delta S^\circ (\text{H}_2)]$$

$$\Delta S^\circ = \left[2 \left(\frac{28.0 \text{ J}}{\text{mol} \cdot \text{K}} \right) + 3 \left(\frac{189.0 \text{ J}}{\text{mol} \cdot \text{K}} \right) \right] - \left[\left(\frac{51.0 \text{ J}}{\text{mol} \cdot \text{K}} \right) + 3 \left(\frac{130.6 \text{ J}}{\text{mol} \cdot \text{K}} \right) \right] = 180.2 \text{ J mol}^{-1} \text{K}^{-1}$$

4. Calculate the standard free energy change, ΔG° , for the complete combustion of methanol, CH₃OH(g), at 25 °C.



compound	ΔG° (kJ mol ⁻¹)
O ₂ (g)	0
H ₂ O(g)	-229.4
CO ₂ (g)	-394.4
CH ₃ OH(g)	-163.2

ANSWER:

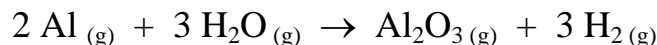
$$\Delta G^\circ = \sum \Delta G^\circ (\text{products}) - \sum \Delta G^\circ (\text{reactants})$$

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

$$\Delta G^\circ = \left[2 \left(\frac{-394.4 \text{ kJ}}{\text{mol}} \right) + 4 \left(\frac{-229.4 \text{ kJ}}{\text{mol}} \right) \right] - \left[2 \left(\frac{-163.2 \text{ kJ}}{\text{mol}} \right) + 3 \left(\frac{0 \text{ kJ}}{\text{mol}} \right) \right] = -1380 \text{ kJ mol}^{-1}$$

$$\Delta G^\circ < 0, \quad \text{The reaction is spontaneous.}$$

5. Calculate the standard free energy change, ΔG° , for the following reaction below, at 25 °C. Also, is this reaction spontaneous?



compound	ΔG° (kJ mol ⁻¹)
Al ₂ O ₃ (g)	-159.0
H ₂ O(g)	-228.6
H ₂ (g)	0
Al(g)	289.4

ANSWER:

$$\Delta G^\circ = \sum \Delta G^\circ (\text{products}) - \sum \Delta G^\circ (\text{reactants})$$

$$\Delta G^\circ = [\Delta G^\circ (\text{Al}_2\text{O}_3) + 3 \Delta G^\circ (\text{H}_2)] - [2 \Delta G^\circ (\text{Al}) + 3 \Delta G^\circ (\text{H}_2\text{O})]$$

$$\Delta G^\circ = \left[\left(\frac{-159.0 \text{ kJ}}{\text{mol}} \right) + 3 \left(\frac{0 \text{ kJ}}{\text{mol}} \right) \right] - \left[2 \left(\frac{289.4 \text{ kJ}}{\text{mol}} \right) + 3 \left(\frac{-228.6 \text{ kJ}}{\text{mol}} \right) \right] = -52.0 \text{ kJ mol}^{-1}$$

$$\Delta G^\circ < 0, \quad \text{The reaction is spontaneous.}$$