

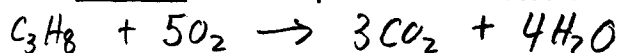
NAME: _____

General Chemistry
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Thermodynamics Calculations, Part II

Solve the following problems. Show formulas, units, and work for problems requiring calculations.

(g). Suppose that 1.00 mol of $C_3H_8(g)$ reacts completely with $O_2(g)$ to produce $CO_2(g)$ and $H_2O(g)$.
Reactants
Products

1. Write the balanced chemical equation for this reaction.



2. Calculate the enthalpy of this reaction. Don't forget the correct units.

$$\text{change } \Delta H_{rxn}^{\circ} = [\text{PRODUCTS}] - [\text{REACTANTS}]$$

$$= [(4)(H_2O) + (3)(CO_2)] - [(1)(C_3H_8) + (5)(O_2)]$$

$$= [(4)(-241.82) + (3)(-393.5)] - [(1)(-103.85) + (5)(0)]$$

$$= [-967.28 + (-1180.5)] - [-103.85]$$

$$= [-2147.78] - [-103.85]$$

$$= -2043.93 \text{ kJ/mol}$$

3. Is this reaction exothermic or endothermic? How do you know?

EXOTHERMIC, because the sign of ΔH is -.

4. Is this enthalpy change favorable or unfavorable?

FAVORABLE.

5. Calculate the ΔS° for the reaction in #1. Use the correct units.

$$\Delta S_{rxn}^{\circ} = [\text{PRODUCTS}] - [\text{REACTANTS}]$$

$$\Delta S_{rxn}^{\circ} = [(4)(H_2O) + (3)(CO_2)] - [(1)(C_3H_8) + (5)(O_2)]$$

$$= [(4)(188.83) + (3)(213.6)] - [(1)(269.9) + (5)(205.0)]$$

$$= [755.32 + 640.8] - [269.9 + 1025]$$

$$= 1396.12 - 1294.9$$

$$= 101.22 \frac{J}{mol \cdot Kelvin}$$

6. Is this entropy change favorable or unfavorable? How do you know?

Favorable. When ΔS_{rxn} is a positive number, the entropy change is favorable.

7. Calculate the ΔG° for this reaction at 298 K. You may use either of the two formulas to calculate ΔG° . Make sure that you keep track of the units -- ΔS° usually has "J" as part of its units, but ΔH° usually has "kJ"

Option #1: use $\Delta G = \Delta H - T\Delta S$

$$\Delta G_{rxn}^\circ = \left[-2043.93 \frac{\text{kJ}}{\text{mol}} \right] - \left[(298 \text{ Kelvins}) \left(0.10122 \frac{\text{kJ}}{\text{mol} \cdot \text{Kelvin}} \right) \right]$$

↑
ANSWER FROM QUESTION #2

↑
GIVEN ABOVE IN THIS PROBLEM

↑
ANSWER FROM #5, BUT IT HAD TO BE CONVERTED TO KJ TO MATCH THE UNITS OF ΔH .

$$101.22 \frac{\text{J}}{\text{mol} \cdot \text{K}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 0.10122 \frac{\text{kJ}}{\text{mol} \cdot \text{K}}$$

$$\Delta G_{rxn}^\circ = \left[-2043.93 \frac{\text{kJ}}{\text{mol}} \right] - \left[30.16 \frac{\text{kJ}}{\text{mol}} \right]$$

$$= -2074.09 \frac{\text{kJ}}{\text{mol}}$$



Option #2: use the "products - reactants" method again.

$$\begin{aligned} \Delta G_{rxn}^\circ &= [\text{PRODUCTS}] - [\text{REACTANTS}] \\ &= [(4)(\text{H}_2\text{O}) + (3)(\text{CO}_2)] - [(1)(\text{C}_3\text{H}_8) + (5)(\text{O}_2)] \\ &= [(4)(-228.57) + (3)(-394.4)] - [(1)(-23.47) + (5)(0)] \\ &= [-914.28 + -1183.2] - [-23.47] \\ &= [-2097.48] - [-23.47] \\ &= -2074.01 \frac{\text{kJ}}{\text{mol}} \end{aligned}$$

← As you can see, you get the same answer using this method (to 4 sig figs, at least).

8. Is the reaction spontaneous or not? How do you know?

It is spontaneous. It releases more energy than it requires.
The sign of ΔG is $-$.

9. Consider the following reaction:



Here is what this expression really means: In this reaction, if 1 mol of X is reacted with 2 mol of W, then 1 mol of Y is produced and this requires 40.9 kJ of energy. Therefore, in order to produce 2.96 mol of Y, how many kJ of energy must be used? Show work. Include units.

$$2.96 \text{ mol of Y} \times \frac{40.9 \text{ kJ}}{1 \text{ mol of X}} = 121.1 \text{ kJ}$$

10. In all of the calculations above (excluding #9), you assumed that 1.00 mol of C_3H_8 reacted. Now, re-calculate the ΔG° from #7 assuming that

- a. Only 0.393 mol of C_3H_8 reacts.

$$0.393 \text{ mol } C_3H_8 \times \frac{-2074 \text{ kJ}}{1 \text{ mol } C_3H_8} = -815.1 \text{ kJ}$$

← answer from # 7

- b. Only 0.393 g of C_3H_8 reacts.

$$0.393 \text{ g} \times \frac{1 \text{ mol}}{44.1 \text{ g}} = 0.008916 \text{ mol} \times \frac{-2074 \text{ kJ}}{1 \text{ mol}} = -18.5 \text{ kJ}$$

$$C_3 = 3 \times 12.0$$

$$H_8 = 8 \times 1.01$$

← use periodic table