

Honors Pre-AP Chemistry—Chapter 6—Thermochemistry Test

Name _____ Date _____ Section _____

Answer the following questions. Questions 1-8 (12 pts each). Question 9 (20 pts).

- 1) If 1.507 g of NH_4Cl is dissolved in 100.0 g of water at 25°C , the temperature of the water drops to 23.98°C . Calculate the heat of the solution (q_{sol}), the heat of the reaction (q_{rxn}), and the ΔH_{rxn} in (kJ/mol). The “s” for water is $4.184 \text{ J/g}^\circ\text{C}$.

- 2) When 57.7 g of copper metal at 99.0°C is added to 100.0 g of water at 27.5°C , the final temperature is found to be 31.0°C . What is the molar heat capacity “s” of the metal? The “s” of the water is $4.184 \text{ J/g}^\circ\text{C}$.

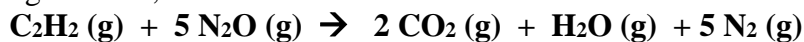
- 3) A 2.02 g sample of quinine, $\text{C}_6\text{H}_4\text{O}_2$, is burned in a bomb calorimeter whose heat capacity “ C_p ” is $7.845 \text{ kJ/}^\circ\text{C}$. The temperature of the calorimeter increases from 23.44°C to 30.57°C . What is the heat of combustion per gram of the quinine? Per mole of quinine?

4) For the following reaction, $2 \text{C(s)} + \text{H}_2 \text{(g)} \rightarrow \text{C}_2\text{H}_2 \text{(g)}$

Apply Hess's Law to the following steps for this reaction.



5) For the following reaction,



Apply Hess's Law to the following steps for this reaction.

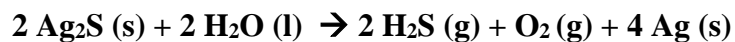


6) Consider the reaction between solid silver sulfide and liquid water producing silver metal, hydrogen sulfide gas, and oxygen gas. **Note:** ΔH_f of $\text{Ag}_2\text{S} = -31.8 \text{ kJ/mol}$.

$$\text{H}_2\text{O} = -241.8 \text{ kJ/mol}$$

$$\text{H}_2\text{S} = -20.2 \text{ kJ/mol}$$

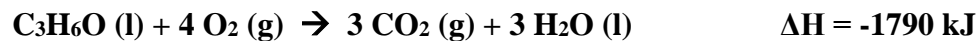
$$\text{Ag} = 0$$



What is the enthalpy of the reaction?

What is the heat required to produce 0.0928 mol of silver? (EC)

7) The complete combustion of acetone, C₃H₆O, results in the liberation of 1790 kJ.



ΔH_f for C₃H₆O = ?

O₂ = 0

CO₂ = -393.5 kJ/mol

H₂O = -285.8 kJ/mol

8) What is the ΔE for a reaction (system) if the surroundings transfer **375 J of heat** to the system and the system exerts **2.3 atm of pressure** on the surroundings changing the volume by **5.2 L**?
Note: (101.3 J/atm L)

- 9) **The standard molar enthalpies of formation for CH₄ (g) and C₂H₆ (g) as follows: -74.8 kJ/mol, and -84.68 kJ/mol respectively.**

Calculate the heat of combustion (ΔH_{rxn}) for these hydrocarbons when they combine with gaseous oxygen (O₂) to produce CO₂ (g) and H₂O (l).

$$\Delta H_f \text{ CO}_2 \text{ (g)} = -393.5 \text{ kJ/mol}$$

$$\Delta H_f \text{ H}_2\text{O (l)} = -285.8 \text{ kJ/mol}$$

What is the $\Delta H/\text{mol}$ for **each** hydrocarbon reaction. Which one is the **more efficient per mole?** (Hint: after calculating ΔH of the reactions, set up ratios to find ΔH for one mole of hydrocarbon)



Equations

$$q = s m \Delta T \quad s_{\text{H}_2\text{O}} = 4.184 \text{ J/g } ^\circ\text{C}$$

$$q_{\text{rxn}} = - q_{\text{sol}} \quad q_{\text{rxn}} = - q_{\text{cal}}$$

$$q_{\text{bomb cal}} = C_p \Delta T$$

$$q_{\text{metal}} = - q_{\text{H}_2\text{O}} \quad \text{or} \quad (m s \Delta T)_{\text{metal}} = - (m s \Delta T)_{\text{H}_2\text{O}}$$

$$\Delta E = q + w \quad w = (P\Delta V)(101.3)$$

$$\Delta H_{\text{rxn}} = \sum n \Delta H_{\text{f}(\text{prod})} - \sum n \Delta H_{\text{f}(\text{react})}$$