

Chapter 17 Thermochemistry Practice Problems Answers

Chapter 17 Thermochemistry Practice Problems Answers Chapter 17 Thermochemistry Practice Problems Answers This blog post provides a comprehensive guide to solving practice problems related to Chapter 17 of a typical chemistry textbook covering the fundamentals of thermochemistry It will delve into the key concepts and formulas required to tackle these problems offering detailed solutions and explanations for each question The aim is to equip students with the necessary tools to understand and apply thermochemical principles effectively Thermochemistry enthalpy entropy Gibbs free energy Hesss Law calorimetry standard enthalpy of formation standard enthalpy of reaction spontaneity equilibrium constant Thermochemistry is a crucial branch of chemistry that deals with the study of heat changes accompanying chemical reactions It explores the relationship between heat flow energy transformations and the chemical and physical properties of substances Chapter 17 of many chemistry textbooks introduces fundamental concepts like enthalpy entropy Gibbs free energy and their role in predicting the spontaneity of reactions This blog post serves as a resource for students to reinforce their understanding of these concepts through the analysis of practice problems Analysis of Current Trends Thermochemistry plays a vital role in various fields including Energy production Understanding energy changes in combustion reactions is crucial for designing efficient power plants and fuel sources Material science Thermodynamic principles guide the development of new materials with desired properties like thermal stability and reactivity Environmental chemistry Assessing the environmental impact of chemical reactions and processes involves understanding heat flow and its impact on ecosystems Biochemistry Thermochemistry is essential for understanding energy transformations within living organisms like cellular respiration and photosynthesis The increasing focus on renewable energy sources sustainable materials and environmental 2 protection underscores the growing relevance of thermochemistry in modern society Discussion of Ethical Considerations Thermochemistry while offering valuable tools for technological advancements also presents ethical considerations Energy consumption The pursuit of energy efficiency often involves the development of new technologies that can have unintended consequences on resource depletion and environmental impact Climate change The

burning of fossil fuels a process governed by thermochemical principles is a significant contributor to greenhouse gas emissions and global warming. Technological development. The advancement of technologies based on thermochemical principles like nuclear power or biofuel production needs to be accompanied by rigorous safety measures and ethical considerations. It is essential to consider the potential ethical ramifications of thermochemical applications and strive for sustainable and responsible practices.

Practice Problems and Solutions

Problem 1 Calculate the enthalpy change for the reaction $2 \text{H}_2\text{g} + \text{O}_2\text{g} \rightarrow 2 \text{H}_2\text{Ol}$. Given the following standard enthalpy of formation values: $\Delta H_f^\circ \text{H}_2\text{Ol} = -285.8 \text{ kJ/mol}$. **Solution** The enthalpy change of a reaction can be calculated using the following equation: $\Delta H = \sum n \Delta H_{f,\text{products}} - \sum m \Delta H_{f,\text{reactants}}$ where ΔH is the enthalpy change of the reaction, ΔH_f° is the standard enthalpy of formation, n and m are the stoichiometric coefficients of the products and reactants respectively. Plugging in the values: $\Delta H = 2(-285.8 \text{ kJ/mol}) - 2(0 \text{ kJ/mol}) - 1(0 \text{ kJ/mol}) = -571.6 \text{ kJ/mol}$. Therefore the enthalpy change for the reaction is -571.6 kJ/mol . This negative value indicates that the reaction is exothermic meaning it releases heat to the surroundings.

Problem 2 A 500 g sample of iron is heated from 250°C to 1000°C . Calculate the heat absorbed by the iron. The specific heat capacity of iron is $0.449 \text{ J/g}^\circ\text{C}$. **Solution** The heat absorbed by a substance can be calculated using the following equation: $q = mC\Delta T$ where q is the heat absorbed, m is the mass of the substance, C is the specific heat capacity, ΔT is the change in temperature. Plugging in the values: $q = 500 \text{ g} \times 0.449 \text{ J/g}^\circ\text{C} \times (1000^\circ\text{C} - 250^\circ\text{C}) = 168375 \text{ J}$. Therefore the heat absorbed by the iron is 168375 J .

Problem 3 A 100 g sample of glucose $\text{C}_6\text{H}_{12}\text{O}_6$ is burned in a calorimeter containing 1000 g of water. The temperature of the water increases from 25.0°C to 27.5°C . Calculate the heat of combustion of glucose in kJ/mol . The specific heat capacity of water is $4.184 \text{ J/g}^\circ\text{C}$. **Solution** First calculate the heat absorbed by the water: $q = 1000 \text{ g} \times 4.184 \text{ J/g}^\circ\text{C} \times (27.5^\circ\text{C} - 25.0^\circ\text{C}) = 10460 \text{ J}$. This heat is released by the combustion of glucose. To find the heat of combustion per mole we need to calculate the moles of glucose burned: $\text{moles of glucose} = \frac{100 \text{ g}}{180.16 \text{ g/mol}} = 0.00555 \text{ mol}$. Therefore the heat of combustion of glucose is $\frac{10460 \text{ J}}{0.00555 \text{ mol}} = 1883720 \text{ J/mol} = 188372 \text{ kJ/mol}$. The heat of combustion of glucose is 188372 kJ/mol .

Problem 4 Using Hess's Law calculate the enthalpy change for the reaction $\text{N}_2\text{g} + 3 \text{H}_2\text{g} \rightarrow 2 \text{NH}_3\text{g}$. Given the following reactions and their enthalpy changes: $\text{N}_2\text{g} + \text{O}_2\text{g} \rightarrow 2 \text{NOg}$, $\Delta H = 180.5 \text{ kJ/mol}$; $2 \text{NOg} + \text{O}_2\text{g} \rightarrow 2 \text{NO}_2\text{g}$, $\Delta H = 114.1 \text{ kJ/mol}$; $4 \text{NH}_3\text{g} + 5 \text{O}_2\text{g} \rightarrow 4 \text{NOg} + 6 \text{H}_2\text{Og}$, $\Delta H = 906.2 \text{ kJ/mol}$; $2 \text{H}_2\text{g} + \text{O}_2\text{g} \rightarrow 2 \text{H}_2\text{Og}$, $\Delta H = 483.6 \text{ kJ/mol}$. **Solution** Hess's Law states that the enthalpy change for a reaction is independent of the pathway taken as long as the initial and final conditions are the same. To calculate the enthalpy change for the target reaction we need to manipulate the given reactions in such a way that they add up to the target reaction.

- Reverse the first reaction: $2 \text{NOg} \rightarrow \text{N}_2\text{g} + \text{O}_2\text{g}$, $\Delta H = -180.5 \text{ kJ/mol}$
- Reverse the second reaction:

$2 \text{ NO}_2\text{g} + 2 \text{ NOg} + \text{O}_2\text{g} \rightarrow \text{H}$ 1141 kJmol
 3 Multiply the third reaction by 12 $2 \text{ NH}_3\text{g} + 5 \text{ O}_2\text{g} + 2 \text{ NOg} + 3 \text{ H}_2\text{Og} \rightarrow \text{H}$ 4531 kJmol
 4 Multiply the fourth reaction by 32 $3 \text{ H}_2\text{g} + 3 \text{ O}_2\text{g} + 3 \text{ H}_2\text{Og} \rightarrow \text{H}$ 7254 kJmol
 5 Add the modified reactions $2 \text{ NOg} + \text{N}_2\text{g} + \text{O}_2\text{g} \rightarrow \text{H}$ 1805 kJmol $2 \text{ NO}_2\text{g} + 2 \text{ NOg} + \text{O}_2\text{g} \rightarrow \text{H}$ 1141 kJmol $2 \text{ NH}_3\text{g} + 5 \text{ O}_2\text{g} + 2 \text{ NOg} + 3 \text{ H}_2\text{Og} \rightarrow \text{H}$ 4531 kJmol $3 \text{ H}_2\text{g} + 3 \text{ O}_2\text{g} + 3 \text{ H}_2\text{Og} \rightarrow \text{H}$ 7254 kJmol $\text{N}_2\text{g} + 3 \text{ H}_2\text{g} + 2 \text{ NH}_3\text{g} \rightarrow \text{H}$ 939 kJmol
 Therefore the enthalpy change for the reaction is 939 kJmol
 Problem 5 Predict whether the following reactions are spontaneous or nonspontaneous at 25 °C
 6 a $2 \text{ NO}_2\text{g} \rightarrow \text{N}_2\text{O}_4\text{g}$ b $\text{CaCO}_3\text{s} \rightarrow \text{CaO}\text{s} + \text{CO}_2\text{g}$
 Given the following standard Gibbs free energy of formation values ΔG_f° NO_2g 513 kJmol ΔG_f° $\text{N}_2\text{O}_4\text{g}$ 979 kJmol ΔG_f° CaCO_3s 11288 kJmol ΔG_f° CaOs 6040 kJmol ΔG_f° CO_2g 3944 kJmol
 Solution The spontaneity of a reaction is determined by the Gibbs free energy change ΔG . If ΔG is negative the reaction is spontaneous and if ΔG is positive the reaction is nonspontaneous.
 a For the reaction $2 \text{ NO}_2\text{g} \rightarrow \text{N}_2\text{O}_4\text{g}$
 $\Delta G = \sum \Delta G_f^\circ(\text{products}) - \sum \Delta G_f^\circ(\text{reactants})$
 $\Delta G = 1(979 \text{ kJmol}) - 2(513 \text{ kJmol}) = -57 \text{ kJmol}$
 Since ΔG is negative the reaction is spontaneous at 25 °C.
 b For the reaction $\text{CaCO}_3\text{s} \rightarrow \text{CaO}\text{s} + \text{CO}_2\text{g}$
 $\Delta G = \sum \Delta G_f^\circ(\text{products}) - \sum \Delta G_f^\circ(\text{reactants})$
 $\Delta G = 1(6040 \text{ kJmol}) + 1(3944 \text{ kJmol}) - 1(11288 \text{ kJmol}) = 1304 \text{ kJmol}$
 Since ΔG is positive the reaction is nonspontaneous at 25 °C.
 Conclusion This blog post has provided a comprehensive overview of thermochemistry covering key concepts and their applications in solving practice problems. By understanding the principles of enthalpy, entropy, Gibbs free energy, and Hess's Law, students can develop a firm grasp of this crucial area of chemistry. While thermochemistry offers powerful tools for technological advancements, it is equally important to consider its ethical implications and strive for sustainable and responsible applications.

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