

chemistry matter change chapter 10 answer key

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Introduction

Welcome to your ultimate study companion for **chemistry matter change chapter 10 answer key**. Whether you're a high school student tackling the final exams, a college freshman navigating your first chemistry course, or a teacher preparing review materials, this guide is crafted to help you master the concepts, solve practice problems, and reinforce your understanding of matter and change.

Chapter 10 of most "Matter and Change" textbooks focuses on the dynamic nature of chemical reactions, energy transformations, and the laws that govern them. In this article, we'll walk through the core ideas, present a curated answer key for the most common exercises, and provide step by step explanations that make the material approachable and memorable.

Let's dive in and transform your study routine into a powerful learning experience!

Overview of Chapter 10: Matter in Motion

Chapter 10 typically covers:

- The distinction between endothermic and exothermic reactions
- Energy profiles and reaction coordinates
- The concept of activation energy and catalysts
- Thermodynamics: enthalpy, entropy, and Gibbs free energy
- Redox reactions and oxidation numbers
- Balancing chemical equations with stoichiometric calculations
- Applications: batteries, combustion, and biochemical pathways

Understanding these topics is essential for answering exam questions that test both conceptual knowledge and quantitative skills. The answer key below is organized by theme, providing not just the correct answers but also the reasoning behind each solution.

Key Concepts Explained

Endothermic vs. Exothermic Reactions

Endothermic reactions absorb heat from the surroundings, resulting in a negative heat flow into the system. Exothermic reactions release heat, giving a positive heat flow out of the system. The sign convention is crucial when interpreting enthalpy changes (ΔH).

Activation Energy and Catalysts

The activation energy (E_a) is the minimum energy barrier that reactants must overcome to form products. Catalysts lower E_a without being consumed, thereby increasing reaction rates. The Arrhenius equation relates reaction rate constants to temperature and E_a .

Thermodynamic Parameters

- Enthalpy (ΔH): Heat change at constant pressure.

- Entropy (9E2“ Measure of disorder.
- Gibbs Free Energy (9Dr“ Determines spontaneity: 9Dr Ò "H – T9E0.

Redox Chemistry

Redox reactions involve the transfer of electrons. Oxidation numbers help track electron flow. The half reaction method balances redox equations by separating oxidation and reduction processes.

Stoichiometry and Balancing Equations

Accurate balancing ensures mass conservation and is the foundation for quantitative predictions in chemical reactions.

Common Practice Questions

Below are representative questions that students frequently encounter. For each, we provide the answer and a concise explanation.

Question 1: Identify the Reaction Type

Which of the following reactions is exothermic?

1. $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$
2. $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$
3. $\text{NaCl}(\text{aq}) + \text{AgNO}_3(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{NaNO}_3(\text{aq})$

Answer: 2. The combustion of hydrogen is a classic exothermic reaction.

Question 2: Calculate ΔH° for a Reaction

Given the standard enthalpies of formation: $\Delta H_f^\circ = -241.8 \text{ kJ/mol}$, $\Delta H_f^\circ = 0$, $\Delta H_f^\circ = 0$, calculate ΔH° for the reaction $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$.

Answer: $\Delta H^\circ = 2\Delta H_f^\circ(\text{H}_2\text{O}) - 2\Delta H_f^\circ(\text{H}_2) - \Delta H_f^\circ(\text{O}_2) = 2(-241.8) - 0 - 0 = -483.6 \text{ kJ}$.

Question 3: Determine Gibbs Free Energy

For the reaction $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$, $\Delta G_f^\circ = -16.5 \text{ kJ/mol}$, $\Delta G_f^\circ = 0$, $\Delta G_f^\circ = 0$. At 298 K, is the reaction spontaneous?

Answer: $\Delta G^\circ = \Delta G_f^\circ(2\text{NH}_3) - \Delta G_f^\circ(\text{N}_2) - 3\Delta G_f^\circ(3\text{H}_2) = 2(-16.5) - 0 - 0 = -33 \text{ kJ}$. The negative ΔG° indicates non-spontaneity at 298 K.

Question 4: Balance a Redox Equation

Balance the following redox reaction in acidic solution: $\text{Cr}_2\text{O}_7^{2-} + 6\text{H}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 3\text{H}_2\text{O}$.

Answer: The equation is already balanced. Key steps: assign oxidation states, confirm electron count, and ensure mass and charge balance.

Question 5: Catalyst Effect on Rate

Which statement best describes the effect of a catalyst on the activation energy of a reaction?

1. Increases E_a
2. Decreases E_a
3. Has no effect on E_a
4. Changes E_a only at high temperatures

Answer: 2. A catalyst lowers the activation energy, thereby increasing the reaction rate.

Answer Key with Detailed Explanations

Below is an expanded answer key for the most common Chapter 10 exercises. Each answer is followed by a brief reasoning section that explains the underlying principles.

Exercise Set 1: Enthalpy Calculations

1. 9D, `or the reaction $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$ is +178 kJ (endothermic). The positive sign reflects energy absorption.
2. 9D, `or the combustion of methane $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$ is -890 kJ (exothermic).
3. 9D, `or the formation of $\text{NaOH}(\text{aq})$ from $\text{Na}(\text{s}) + \frac{1}{2}\text{O}_2(\text{g}) + \frac{1}{2}\text{H}_2\text{O}(\text{l})$ is -469 kJ.

Key Insight: Always use standard enthalpies of formation; remember that $\Delta H^\circ_{\text{f}}(\text{O}_2) = 0$ and $\Delta H^\circ_{\text{f}}(\text{H}_2\text{O}) = -285.8 \text{ kJ/mol}$.

Exercise Set 2: Gibbs Free Energy and Spontaneity

1. For the reaction $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$ at 298 K, $\Delta G^\circ < 0$, indicating a spontaneous process.
2. For the reaction $\text{C}_2\text{H}_6(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightarrow \text{C}_2\text{H}_5\text{OH}(\text{g})$ at 298 K, $\Delta G^\circ > 0$; the reaction is non-spontaneous under standard conditions.
3. For the reaction $\text{Fe}(\text{s}) + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Fe}^{2+}(\text{aq}) + \text{Cu}(\text{s})$ at 298 K, $\Delta G^\circ < 0$, spontaneous.

Key Insight: A negative ΔG° indicates spontaneity at constant temperature and pressure.

Exercise Set 3: Redox Balancing

1. Oxidation: $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$; Reduction: $\text{Cu}^{2+} + \text{e}^- \rightarrow \text{Cu}(\text{s})$. Combined: $\text{Fe}^{2+} + \text{Cu}^{2+} \rightarrow \text{Fe}^{3+} + \text{Cu}(\text{s})$.
2. Oxidation: $\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$; Reduction: $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$. Multiply the half-reactions to equalize electrons and combine.
3. Oxidation: $\text{H}_2 \rightarrow 2\text{H}^+ + 2\text{e}^-$; Reduction: $\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}$. Multiply the hydrogen half-reaction by 2.

Key Insight: Always balance atoms first, then charge by adding electrons, and finally combine the half-reactions.

Exercise Set 4: Catalysis and Kinetics

1. The reaction rate increases by a factor of 10 when a platinum catalyst is introduced.
2. The activation energy of the decomposition of hydrogen peroxide in the presence of potassium iodide is reduced from 75 kJ to 35 kJ.
3. In the presence of a catalyst, the reaction order remains unchanged but the rate constant increases.

Key Insight: Catalysts lower E_a , which, per the Arrhenius equation, leads to a higher rate constant ($k = A e^{(-E_a/RT)}$).

How to Use This Answer Key Effectively

1. **Practice First, Then Verify:** Attempt each problem without looking at the answer. Only after you've given it a full effort should you consult the key.
2. **Review the Explanations:** The reasoning behind each solution is as valuable as the answer itself. Understanding the logic helps you tackle novel problems.
3. **Create Flashcards:** Turn key concepts (e.g., $\Delta H^\circ_{\text{f}}(\text{O}_2) = 0$) into flashcards for quick recall.
4. **Simulate Exam Conditions:** Time yourself while solving the exercises to build test-taking stamina.
5. **Teach Someone Else:** Explaining the solution to a peer reinforces your own understanding.

Study Tips for Chapter 10 Mastery

- Concept Mapping: Draw a diagram linking endothermic/exothermic, ΔH° , ΔG° , ΔS° —G•.
- Use Real World Examples: Relate combustion to car engines, redox to batteries, and catalysis to industrial processes.
- Work on Word Problems: Practice translating natural language scenarios into equations.
- Review Past Exams: Identify recurring question types and focus on those areas.
- Group Study: Discuss challenging problems with classmates to gain multiple perspectives.

Frequently Asked Questions

What if the answer key is missing a step?

When a solution omits intermediate steps, recreate the derivation yourself. This reinforces the procedural knowledge required for exams.

How do I know if a reaction is spontaneous?

Calculate $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$; if $\Delta G^\circ < 0$

Can a catalyst reverse a reaction's direction?

No. A catalyst speeds up both the forward and reverse reactions equally, shifting the equilibrium position only if the concentrations of reactants or products change.

Why is the standard enthalpy of formation for elements zero?

By definition, the standard state of an element is its most stable form at 1 atm and 25 °C. Therefore, its formation enthalpy is set to zero as a reference point.

Is it always safe to use the Arrhenius equation?

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