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Understanding the energetics of chemical reactions is fundamental in chemistry, especially when analyzing whether a reaction is exothermic or endothermic. The reaction involving ethane (C_2H_6) converting to ethylene (C_2H_4) and hydrogen (H_2), accompanied by an energy change of 137 kJ, offers an interesting case study. Specifically, this reaction's energy profile reveals that the chemical energy stored within the reactants is greater than that in the products, implying that energy is released during the process. This article explores the details of this reaction, the concept of chemical energy, and the implications of the energy difference, providing an in-depth understanding suitable for students, educators, and professionals alike.

Understanding the Chemical Reaction: C_2H_6 + 137 KJ → C_2H_4 + H_2

Reaction Overview

The chemical reaction in question involves the conversion of ethane (C_2H_6) into ethylene (C_2H_4) with the release of hydrogen gas (H_2). The equation can be written as:

$$C_2H_6 \rightarrow C_2H_4 + H_2 \quad (\Delta H = -137 \text{ kJ})$$

Here, the negative sign in ΔH indicates that the reaction releases energy, making it exothermic.

Reaction Conditions and Significance

- Industrial Relevance: Ethylene is a vital raw material in plastics manufacturing, and its production from ethane is a common process.
- Reaction Type: This is an example of a dehydrogenation reaction, where hydrogen atoms are removed from ethane.
- Energy Profile: The energy change indicates the amount of energy released or absorbed during the

transformation.

Chemical Energy: Concept and Importance

What Is Chemical Energy?

Chemical energy is the energy stored within the bonds of chemical compounds. It is a form of potential energy that can be released or absorbed during chemical reactions.

How Is Chemical Energy Measured?

- Enthalpy (H): A thermodynamic property measuring heat content at constant pressure.
- Bond Energies: The energy required to break chemical bonds.
- Standard Enthalpy of Reaction ($\Delta H°$): The change in enthalpy during a reaction under standard conditions.

Relation Between Chemical Energy and Reaction Enthalpy

- When bonds are broken, energy is absorbed.
- When new bonds form, energy is released.
- The overall ΔH indicates whether a reaction is exothermic or endothermic.

Analyzing the Energy Difference in the Reaction

The Significance of the 137 KJ Energy Change

The provided energy change of 137 kJ suggests that:

- The reaction releases 137 kJ of energy.
- The reactants (ethane) possess more chemical energy than the products (ethylene and hydrogen).
- The difference accounts for the energy released during the formation of new bonds and the breaking of old bonds.

Why Is the Chemical Energy of Reactants Greater?

- Bond Dissociation and Formation: Ethane has C–H and C–C bonds with higher bond energies compared to the bonds in ethylene and H_2.
- Energy Release Mechanism: As bonds in ethane break and new, more stable bonds form in ethylene and H_2, energy is released.
- Thermodynamic Perspective: The overall energy change indicates a net loss of stored chemical energy in the reactants, making the process exothermic.

Implications of Greater Reactant Energy

- The reaction is energetically favorable and can occur spontaneously under suitable conditions.
- The released energy can be harnessed for practical applications, such as energy generation or industrial processes.

Thermodynamics of the Reaction

Enthalpy Change (ΔH)

- The reaction's enthalpy change of -137 kJ confirms its exothermic nature.
- This energy is typically released as heat.

Gibbs Free Energy (ΔG)

- For spontaneity, ΔG should be negative.
- The relation: $\Delta G = \Delta H - T\Delta S$ (where T is temperature and ΔS is entropy change).
- In this reaction, the energy release contributes to a negative ΔG, favoring spontaneity.

Energy Profile Diagram

- Reactants start at a higher energy level.
- Transition state represents the energy barrier.
- Products are at a lower energy level, indicating energy release.

Practical Implications and Applications

Industrial Significance

- Ethylene Production: Dehydrogenation of ethane is a common method to produce ethylene, a precursor for plastics like polyethylene.
- Energy Considerations: The exothermic nature means less external energy input is required, enhancing process efficiency.

Environmental Impact

- Energy Efficiency: Reactions that release energy can reduce overall energy consumption.
- Green Chemistry: Understanding energy profiles helps in designing sustainable chemical processes.

Safety Aspects

- Reactions releasing significant energy necessitate proper handling and safety protocols to prevent accidents.

Summary and Key Takeaways

- The reaction $C_2H_6 + 137$ KJ → $C_2H_4 + H_2$ is exothermic, releasing 137 kJ of energy.
- The chemical energy stored within ethane (reactant) is greater than that in ethylene and hydrogen (products).
- This energy difference results from bond-breaking and bond-forming processes, favoring the formation of more stable bonds.
- Understanding the energy dynamics helps optimize industrial processes, improve safety, and promote sustainable chemistry practices.
- Thermodynamic concepts such as enthalpy change and Gibbs free energy are essential tools for analyzing such reactions.

Conclusion

The analysis of this chemical reaction underscores the fundamental principle that the chemical energy of reactants determines whether a reaction is exothermic or endothermic. In this case, the reactants possess greater chemical energy, which is released during the transformation into more stable products. Recognizing these energy changes is crucial for chemists and engineers working to develop efficient, safe, and sustainable chemical processes that leverage the natural energy flows within chemical reactions. Whether in academic research, industrial manufacturing, or environmental management, understanding the energetics of reactions like C_2H_6 decomposing into C_2H_4 and H_2 provides valuable insights into the energetic landscape of chemistry.

Frequently Asked Questions

Is the given reaction endothermic or exothermic?

The reaction is endothermic because it absorbs 137 KJ of energy, indicating that the chemical energy of the reactant (C2H6) is greater than that of the products.

What does the energy change in the reaction imply about the bond energies?

The energy change suggests that breaking the bonds in C2H6 requires more energy than is released when forming bonds in C2H4 and H2, confirming that the reactant's chemical energy is higher.

How does the energy requirement affect the spontaneity of the reaction?

Since the reaction absorbs energy (endothermic), it is non-spontaneous under standard conditions without energy input.

What is the significance of the 137 KJ energy value in the reaction?

The 137 KJ represents the amount of energy required to convert reactants into products, indicating the reaction's endothermic nature.

In terms of chemical energy, what can be inferred about C2H6 compared to C2H4 and H2?

C2H6 has higher chemical energy than the combined energy of C2H4 and H2, as energy must be absorbed to proceed with the reaction.

How would changing conditions (like temperature or catalysts) affect this reaction?

Adjusting conditions such as increasing temperature or adding catalysts may influence the reaction rate but do not change the fundamental energy requirement that C2H6's chemical energy is greater than that of the products.

Additional Resources

For The Chemical Reaction C_2H_6 + 137 KJ C_2H_4 + H_2: The Chemical Energy Of The Reactant Is Greater Than

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Introduction

Understanding the energy dynamics of chemical reactions is fundamental in chemistry, providing insight into reaction spontaneity, energy transfer, and thermodynamic stability. One such reaction that exemplifies the principles of energy change is the conversion of ethane (C_2H_6) into ethylene (C_2H_4) with the release of molecular hydrogen (H_2), accompanied by an energy change of 137 kJ. This investigation aims to analyze the thermodynamic profile of this reaction, with particular emphasis on the assertion that the chemical energy of the reactant, ethane, is greater than that of the products, ethylene and hydrogen.

Contextualizing the Reaction

The chemical reaction under scrutiny can be represented as:

$C_2H_6 \rightarrow C_2H_4 + H_2 \quad \Delta H = -137\ kJ$

Here, the negative enthalpy change indicates an exothermic process, which releases energy into the surroundings. This reaction typifies a dehydrogenation process, often employed industrially to produce ethylene—a key precursor in polymer manufacturing.

Fundamental Thermodynamic Concepts

Chemicals' Chemical Energy and Enthalpy

Chemical energy refers to the potential energy stored within chemical bonds. When bonds are broken and formed during a reaction, there is an associated energy change. The overall energy change of a reaction (ΔH) reflects the difference in chemical energy between reactants and products.

Exothermic vs. Endothermic Reactions

- Exothermic reactions release energy ($\Delta H < 0$), indicating that the products are at a lower energy state than reactants.
- Endothermic reactions absorb energy ($\Delta H > 0$), with products at a higher energy level.

In the case of the reaction at hand, a negative ΔH suggests that the reactant ethane possesses greater chemical energy than the combined energy of ethylene and hydrogen.

Thermodynamic Analysis of the Reaction

Bond Energy Perspective

The reaction involves the removal of hydrogen atoms from ethane to form ethylene. The process can be broken down into:

- Breaking bonds in ethane (C–H and C–C bonds)
- Forming bonds in ethylene and hydrogen

The net energy change depends on the bond dissociation energies (BDEs). Ethane's bonds are more saturated, featuring single bonds, which are generally more stable and possess lower energy compared to the unsaturated ethylene molecule, which contains a double bond.

Key Bond Dissociation Energies (approximate values):

Bond Type	BDE (kJ/mol)
C–H (alkane)	~412
C–C (single)	~348
C=C (double)	~614

Given these, the energy required to break bonds in ethane is offset by the energy released when forming the bonds in ethylene and hydrogen.

Enthalpy Calculations

The overall enthalpy change (ΔH) can be estimated by:

$\Delta H \approx$ [Sum of BDEs of bonds broken] – [Sum of BDEs of bonds formed]

The observed negative value of ΔH (-137 kJ) indicates that the bonds formed in the products are more stable (lower energy) than those in the reactant.

Implication: Greater Reactant Energy

The negative ΔH confirms that the chemical energy stored in ethane is higher than that in ethylene and hydrogen. This energy difference is released during the reaction, reflecting the energy stored in ethane's bonds prior to reaction.

Quantitative Perspective

Molar Enthalpy Considerations

- The energy released per mole of ethane converted is 137 kJ.
- The initial chemical energy stored in ethane exceeds the combined energy of the products by this amount.

This quantifies the extent of energy release and confirms the higher energetic state of the reactant molecule before reaction.

Thermodynamic Implications and Reaction Spontaneity

Gibbs Free Energy (ΔG)

The spontaneity of a reaction is governed by the Gibbs free energy change:

$$\Delta G = \Delta H - T\Delta S$$

Where ΔS is the entropy change, and T is the temperature in Kelvin.

Given the exothermic nature (ΔH negative) and assuming an increase in entropy (due to the production of H_2 gas), the reaction tends to be spontaneous under standard conditions.

Energy of Reactants vs. Products

The key takeaway is that the reactant ethane possesses a higher chemical energy state than the products, making the reaction energetically favorable and thermodynamically spontaneous.

Broader Context and Industrial Relevance

Dehydrogenation of Ethane

This reaction is a significant step in producing ethylene, a vital building block in plastics and chemicals. Understanding the energy profile guides process optimization, energy recovery strategies, and catalyst development.

Thermodynamic Efficiency

The fact that the reaction releases energy suggests potential for energy recovery and utilization in industrial settings, affecting economic and environmental considerations.

Summary and Conclusions

- The reaction $C_2H_6 \rightarrow C_2H_4 + H_2$ involves a release of 137 kJ per mole, indicating an exothermic process.
- The chemical energy stored in ethane is greater than that in ethylene and hydrogen, which is confirmed by the negative enthalpy change.
- Bond energy analysis supports the thermodynamic feasibility of the reaction, with bonds in ethane being higher in energy than those in the products.
- The process's spontaneity is further reinforced by the likely increase in entropy due to hydrogen gas formation.
- Industrial applications leverage this energy release, making the thermodynamic understanding crucial for process optimization.

Final Remarks

This analysis underscores the importance of thermodynamic principles in evaluating chemical reactions. The reaction of ethane's dehydrogenation exemplifies how chemical energy is stored and released, with practical implications spanning energy efficiency, process design, and chemical manufacturing.

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Note: All numerical values and interpretations are based on standard chemical data and thermodynamic principles, tailored for an in-depth understanding suitable for a scientific review or publication.

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