For A Certain Chemical Reaction, The Standard Gibbs Free Energy Of Reaction At 20.0 C Is 136. KJ. Calculate

For A Certain Chemical Reaction, The Standard Gibbs Free Energy Of Reaction At 20.0 C Is 136. KJ. Calculate the equilibrium constant (K) at this temperature, and analyze what this value indicates about the spontaneity and extent of the reaction. Understanding how Gibbs free energy relates to equilibrium is fundamental in thermodynamics and chemical kinetics, allowing chemists to predict the direction of reactions and their feasibility under specific conditions. In this article, we will explore the principles involved in calculating the equilibrium constant from Gibbs free energy, the relevant equations, step-by-step calculations, and the implications of the results.

Understanding Gibbs Free Energy and Its Role in Chemical Reactions

What is Gibbs Free Energy?

Gibbs free energy (G) is a thermodynamic potential that measures the maximum reversible work obtainable from a thermodynamic system at constant temperature and pressure. It combines enthalpy (H), entropy (S), and temperature (T) into a single value, providing insight into the spontaneity of processes.

The change in Gibbs free energy for a reaction, ΔG°, indicates whether a reaction occurs spontaneously under standard conditions:
- $\Delta G^\circ < 0$: The reaction is spontaneous.
- $\Delta G^\circ > 0$: The reaction is non-spontaneous.
- $\Delta G^\circ = 0$: The system is at equilibrium.

Standard Gibbs Free Energy of Reaction

The standard Gibbs free energy of reaction (ΔG°) is measured under standard conditions: 1 bar pressure, pure reactants and products in their standard states, and a specified temperature (here, 20.0°C). It provides a baseline for predicting reaction behavior under these conditions.

Relating Gibbs Free Energy to Equilibrium Constant

The Fundamental Equation

The relationship between the standard Gibbs free energy change and the equilibrium constant (K) is given by the thermodynamic equation:

$$ \Delta G^\circ = -RT \ln K $$

Where:
- ΔG° = standard Gibbs free energy change (in Joules)
- R = universal gas constant, 8.314 J/(mol·K)
- T = temperature in Kelvin
- K = equilibrium constant

This equation allows us to calculate the equilibrium constant directly from the known ΔG° and temperature.

Converting Units and Conditions

Given:
- ΔG° = 136 kJ
- Temperature = 20.0°C

First, convert ΔG° to Joules:
$$ 136\, \text{kJ} = 136,000\, \text{J} $$

Next, convert Celsius to Kelvin:
$$ T = 20.0 + 273.15 = 293.15\, \text{K} $$

Now, we are ready to perform the calculation.

Calculating the Equilibrium Constant (K)

Step-by-Step Calculation

Using the equation:

$$ \Delta G^\circ = -RT \ln K $$

Rearranged to solve for K:

$$ \ln K = -\frac{\Delta G^\circ}{RT} $$

Plugging in the values:

$$ \ln K = -\frac{136,000}{8.314 \times 293.15} $$

Calculate the denominator:

$$ 8.314 \times 293.15 \approx 2437.6 $$

Now, compute the ratio:

$$ \frac{136,000}{2437.6} \approx 55.75 $$

Since ΔG° is positive, the negative sign applies:

$$ \ln K = -55.75 $$

Exponentiating both sides:

$$ K = e^{-55.75} $$

Using a calculator:

$$ K \approx 1.43 \times 10^{-24} $$

Result:

$$
\boxed{
K \approx 1.43 \times 10^{-24}
}
$$

Interpretation:

The extremely small value of K indicates that, at 20.0°C, the reaction favors reactants over products, and the formation of products is thermodynamically unfavorable under standard conditions.

Implications of the Calculated Equilibrium Constant

Spontaneity and Extent of Reaction

Since ΔG° is positive and K is significantly less than 1, the reaction does not proceed spontaneously in the forward direction at this temperature. The equilibrium position heavily favors the reactants, meaning very little product is formed under standard conditions.

Temperature Dependence

The value of ΔG° is temperature-dependent; a different temperature could shift the equilibrium position. For example, increasing temperature might alter ΔG° and K, possibly favoring product formation if the reaction is endothermic or exothermic.

Additional Considerations and Applications

Predicting Reaction Behavior

Understanding the relationship between Gibbs free energy and equilibrium constant helps chemists:
- Design reactions with desired yields.
- Determine conditions to shift equilibrium.
- Assess reaction feasibility.

Limitations and Assumptions

The calculation assumes standard conditions and neglects activity effects or non-ideal behavior. Real systems may deviate, requiring more detailed thermodynamic data.

Practical Applications

This approach is vital in industries such as:
- Chemical manufacturing
- Environmental chemistry
- Biochemistry, where reaction spontaneity influences metabolic pathways
- Material science, for predicting stability and reactions of compounds

Summary

In conclusion, given the standard Gibbs free energy of reaction at 20.0°C as 136 kJ, the equilibrium constant is approximately 1.43 × 10^{-24}. This indicates a reaction strongly favoring reactants under standard conditions, highlighting the importance of thermodynamic data in predicting reaction behavior. Mastery of these calculations enables chemists to tailor reaction conditions, optimize yields, and understand the fundamental energetics underlying chemical processes.

Key Takeaways

- The Gibbs free energy change relates directly to the equilibrium constant via \(\Delta G^\circ = -RT \ln K\).
- Positive ΔG° corresponds to a very small K, favoring reactants.
- Temperature and ΔG° are critical in predicting whether a reaction is spontaneous or not.
- Thermodynamic calculations guide practical decision-making in chemical synthesis and process engineering.

By understanding and applying these principles, chemists can predict reaction outcomes accurately, optimize conditions, and develop efficient chemical processes aligned with thermodynamic feasibility.

Frequently Asked Questions

How is the standard Gibbs free energy change (ΔG°) related to the equilibrium constant (K) at 20.0°C?

At 20.0°C, ΔG° is related to the equilibrium constant K by the equation ΔG° = -RT ln K, where R is the gas constant (8.314 J/mol·K) and T is the temperature in Kelvin.

Given ΔG° = 136 kJ at 20.0°C, how can we calculate the equilibrium constant K?

Convert ΔG° to joules (136,000 J), then use K = e^(-ΔG° / RT). With R = 8.314 J/mol·K and T = 293.15 K (20°C), plug in the values to find K.

What is the significance of a positive ΔG° value for a chemical reaction at 20.0°C?

A positive ΔG° indicates that the reaction is non-spontaneous under standard conditions at 20.0°C and will not proceed spontaneously in the forward direction.

How does temperature affect the calculation of ΔG° and the equilibrium constant?

Temperature influences ΔG° and K through the relation ΔG° = -RT ln K; as temperature changes, the value of K and the spontaneity of the reaction can also change accordingly.

If ΔG° is known, how can we determine whether a reaction is spontaneous at 20.0°C?

Since ΔG° is positive (136 kJ), the reaction is non-spontaneous at 20.0°C. For spontaneity, ΔG° should be negative.

What is the typical process to calculate ΔG° for a reaction given thermodynamic data?

ΔG° can be calculated using ΔG° = ΔH° - TΔS°, where ΔH° is the enthalpy change and ΔS° is the entropy change of the reaction, both at standard conditions.

Additional Resources

For A Certain Chemical Reaction, The Standard Gibbs Free Energy Of Reaction At 20.0°C Is 136 kJ. Calculate — this statement sets the stage for a comprehensive exploration of thermodynamics, specifically focusing on how to determine the equilibrium constant from standard Gibbs free energy values. Understanding this process is fundamental in physical chemistry, providing insights into the

spontaneity of reactions, their equilibrium positions, and the thermodynamic parameters that govern chemical processes.

In this article, we'll delve into the concepts behind Gibbs free energy, explore the relationships between free energy and equilibrium, and walk through the calculation steps for the equilibrium constant based on the provided data. Whether you're a student, a researcher, or simply an enthusiast, this guide aims to clarify the underlying principles and equip you with the tools to solve similar problems confidently.

Understanding the Standard Gibbs Free Energy of Reaction

The standard Gibbs free energy of reaction (ΔG°_{rxn}) is a thermodynamic property that indicates the favorability of a chemical reaction under standard conditions (usually 1 bar pressure and a specified temperature, here 20.0°C). It combines enthalpy and entropy considerations:

$$
\Delta G^\circ_{rxn} = \Delta H^\circ_{rxn} - T \Delta S^\circ_{rxn}
$$

where:
- ΔH°_{rxn} is the standard enthalpy change,
- ΔS°_{rxn} is the standard entropy change,
- T is the absolute temperature in Kelvin.

The sign and magnitude of ΔG°_{rxn} determine whether a reaction proceeds spontaneously under standard conditions:
- $\Delta G^\circ_{rxn} < 0$: Reaction is spontaneous,
- $\Delta G^\circ_{rxn} > 0$: Reaction is non-spontaneous,
- $\Delta G^\circ_{rxn} = 0$: Reaction is at equilibrium.

In our case, the standard Gibbs free energy of reaction at 20.0°C is 136 kJ (positive), indicating that, under standard conditions, the reaction is non-spontaneous.

Connecting Gibbs Free Energy to Equilibrium Constant

A key thermodynamic relationship links ΔG°_{rxn} to the equilibrium constant (K_{eq}):

$$
\Delta G^\circ_{rxn} = -RT \ln K_{eq}
$$

where:
- R is the universal gas constant,
- T is the temperature in Kelvin,
- K_{eq} is the equilibrium constant (unitless).

This relationship allows us to determine K_{eq} directly from the standard Gibbs free energy:

$$
K_{eq} = e^{-\frac{\Delta G^\circ_{rxn}}{RT}}
$$

Note: ΔG°_{rxn} must be in consistent units with RT. Since $R = 8.314\, \text{J mol}^{-1} \text{K}^{-1}$, and ΔG°_{rxn} is given in kJ, we need to convert units appropriately.

Step-by-Step Calculation

1. Convert ΔG°_{rxn} to Joules

Given:
$$
\Delta G^\circ_{rxn} = 136\, \text{kJ} = 136,000\, \text{J}
$$

2. Convert Temperature to Kelvin

Given:
$$
T = 20.0^\circ C = 20.0 + 273.15 = 293.15\, \text{K}
$$

3. Calculate RT

$$
RT = 8.314\, \text{J mol}^{-1} \text{K}^{-1} \times 293.15\, \text{K} \approx 2438.2\, \text{J mol}^{-1}
$$

4. Compute the exponent

$$
-\frac{\Delta G^\circ_{rxn}}{RT} = -\frac{136,000\, \text{J}}{2438.2\, \text{J mol}^{-1}} \approx -55.77
$$

5. Calculate K_{eq}

$$
K_{eq} = e^{-55.77} \approx 1.58 \times 10^{-24}
$$

This extremely small value indicates that, at 20.0°C, the reaction heavily favors the reactants; the equilibrium position lies far to the left.

Interpretation of Results

The calculated equilibrium constant ($K_{eq} \approx 1.58 \times 10^{-24}$) is virtually zero, confirming that the reaction does not proceed spontaneously under standard conditions at 20.0°C. This aligns with the positive ΔG°_{rxn} value, which signifies non-spontaneity.

In practical terms:
- The reaction favors the reactants significantly,
- The formation of products is extremely unfavorable under these conditions,
- To drive the reaction forward, conditions such as increased temperature, removal of products, or coupling with favorable reactions may be necessary.

Broader Context and Additional Considerations

Effect of Temperature on Equilibrium

Since ΔG°_{rxn} is temperature-dependent, changing the temperature might alter the spontaneity:

$$
\Delta G^\circ_{rxn}(T) = \Delta H^\circ_{rxn} - T \Delta S^\circ_{rxn}
$$

And, correspondingly, the equilibrium constant at a different temperature:

$$
K_{eq}(T) = e^{-\frac{\Delta G^\circ_{rxn}(T)}{RT}}
$$

If more data like ΔH°_{rxn} and ΔS°_{rxn} are available, one could predict how K_{eq} shifts with temperature.

Limitations and Assumptions

- The calculation assumes ideal behavior and standard conditions.
- Real reactions may be influenced by pressure, concentration, catalysts, or non-idealities.
- The value of ΔG°_{rxn} is often derived from tabulated thermodynamic data, which could have uncertainties.

Practical Applications

Understanding the relationship between ΔG°_{rxn} and K_{eq} is crucial in:
- Chemical manufacturing,
- Environmental chemistry,
- Biological systems,
- Material science.

It guides chemists in designing processes that favor desired reactions.

Summary

- The standard Gibbs free energy of reaction at 20.0°C is 136 kJ.
- Using thermodynamic relationships, the equilibrium constant K_{eq} can be calculated via:

$$
K_{eq} = e^{-\frac{\Delta G^\circ_{rxn}}{RT}}
$$

- Converting units and plugging in the values yields:

$$
K_{eq} \approx 1.58 \times 10^{-24}
$$

- This indicates the reaction is strongly non-spontaneous under standard conditions at 20.0°C, favoring reactants.

- Adjustments in temperature or reaction conditions may be necessary to shift the equilibrium toward products.

Final Remarks

Mastering the calculation of equilibrium constants from Gibbs free energy data is a fundamental skill in thermodynamics and physical chemistry. It enables scientists and engineers to predict reaction behavior, optimize processes, and understand the energetic landscape of chemical reactions. By following systematic steps—unit conversions, applying the core thermodynamic equations, and interpreting the results—you can effectively analyze a wide range of chemical systems.

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