A Reaction Has A Standard Freeenergy Change Of 11.60 Kj Mol1(2.772 Kcal Mol1). Calculate The Equilibrium

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Understanding the relationship between standard free energy change and chemical equilibrium is fundamental in chemistry. When a reaction has a standard free energy change (ΔG°) of 11.60 kJ/mol (or 2.772 kcal/mol), it indicates how far the reaction is from equilibrium under standard conditions. Calculating the equilibrium constant (K) from this value allows chemists to predict whether a reaction will favor reactants or products. This article will guide you through the process of calculating the equilibrium constant based on the given ΔG°, explore the significance of these calculations, and provide practical insights into their applications.

Understanding Standard Free Energy Change (▲G°)

What Is ▲G°?

Standard free energy change (ΔG°) measures the spontaneity of a chemical reaction under standard conditions, which are typically 1 bar pressure, 1 M concentration for solutions, and a specified temperature (usually 25°C or 298 K). When ΔG° is negative, the reaction favors the formation of products; when positive, it favors reactants.

The Relationship Between ▲G° and Equilibrium Constant (K)

The key equation connecting free energy and equilibrium is:

$$\Delta G° = -RT \ln K$$

Where:
- $\Delta G°$ = Standard free energy change in joules per mole (J/mol)
- R = Universal gas constant (8.314 J/mol·K)
- T = Temperature in kelvin (K)
- K = Equilibrium constant (unitless)

This equation allows us to calculate K when ΔG° and temperature are known, providing insight into the extent of the reaction at equilibrium.

Calculating the Equilibrium Constant (K)

Step 1: Convert ΔG° to Joules

Given:
- ΔG° = 11.60 kJ/mol

Convert to joules:
$$ 11.60\, \text{kJ/mol} = 11,600\, \text{J/mol} $$

Step 2: Plug Values into the Equation

Assuming standard temperature:
- $ T = 298\, \text{K} $
- $ R = 8.314\, \text{J/mol·K} $

Rearranged equation to solve for $ K $:
$$ K = e^{-\Delta G° / (RT)} $$

Calculate the exponent:
$$ -\frac{\Delta G°}{RT} = -\frac{11,600}{8.314 \times 298} $$

Compute denominator:
$$ 8.314 \times 298 \approx 2477.772 $$

Compute the ratio:
$$ -\frac{11,600}{2477.772} \approx -4.677 $$

Calculate $ K $:
$$ K = e^{-4.677} $$

Using a calculator:
$$ K \approx 0.0093 $$

Step 3: Interpret the Result

The equilibrium constant:
$$ K \approx 0.0093 $$

This value indicates that at equilibrium, the concentration of products is much lower than that of reactants, implying the reaction predominantly favors reactants under standard conditions.

Implications of the Equilibrium Constant

Reaction Favorability

- Since $K < 1$, the reaction favors the reactants.
- The smaller the K, the less the reaction proceeds toward products at equilibrium.

Effect of Temperature

- The calculation assumes standard temperature (298 K). Changes in temperature will alter K.
- For endothermic reactions (positive ΔG°), increasing temperature may shift equilibrium toward products.
- For exothermic reactions (negative ΔG°), increasing temperature favors reactants.

Practical Applications in Chemistry

- Predicting reaction yields in chemical manufacturing.
- Designing processes to maximize product formation.
- Understanding biological pathways and enzyme functions.
- Developing pharmaceuticals and other chemical products.

Additional Considerations in Free Energy and Equilibrium Calculations

Effect of Concentrations and Conditions

While ΔG° provides a baseline under standard conditions, actual reactions often occur under non-standard conditions. The reaction quotient Q (actual concentrations) can be used to determine the direction in which a reaction will proceed.

$$\Delta G = \Delta G° + RT \ln Q$$

- If $\Delta G < 0$, the reaction proceeds forward.
- If $\Delta G > 0$, the reaction proceeds in reverse.
- When $\Delta G = 0$, the system is at equilibrium, and $Q = K$.

Limitations and Assumptions

- The calculation assumes standard conditions and no side reactions.
- Temperature is a critical factor; deviations can significantly impact K.
- Real systems may have kinetic barriers that affect the rate, but ΔG° deals with thermodynamic feasibility.

Summary and Key Takeaways

- Standard free energy change (ΔG°) directly relates to the position of equilibrium via the equation $\Delta G° = -RT \ln K$.
- Given ΔG° = 11.60 kJ/mol, the equilibrium constant K at 298 K is approximately 0.0093, indicating a reaction that favors reactants.
- Understanding K helps chemists predict reaction outcomes, optimize conditions, and design efficient processes.
- Temperature and actual concentrations influence the reaction’s direction and extent, which can be analyzed using the Gibbs free energy equations.
- These calculations are essential in fields ranging from industrial chemistry to biochemistry and pharmaceutical development.

By mastering the relationship between free energy and equilibrium, scientists and engineers can better control chemical reactions, improve yields, and develop innovative solutions across various industries. Whether you're a student learning about thermodynamics or a professional designing reactions, understanding these principles is crucial for success.

Frequently Asked Questions

How do you calculate the equilibrium constant (K) from the standard free energy change (ΔG°)?

You can calculate the equilibrium constant using the formula ΔG° = -RT ln K, where R is the gas constant and T is the temperature in Kelvin. Rearranged, K = e^(-ΔG° / RT).

What is the value of the equilibrium constant for the reaction with ΔG° = 11.60 kJ/mol at 298 K?

Using the formula K = e^(-ΔG° / RT): K = e^(-11600 / (8.314 298)) ≈ e^(-4.66) ≈ 0.0094.

Why is the equilibrium constant less than 1 for this reaction?

Because the positive ΔG° indicates the reaction is non-spontaneous under standard conditions, resulting in an equilibrium constant less than 1, favoring reactants.

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

Temperature directly affects the calculation through the RT term in the exponent. Higher temperatures can increase or decrease K depending on whether ΔG° is positive or negative.

Can we determine the direction of the reaction from the sign of ΔG°? How?

Yes. A positive ΔG° indicates the reaction favors reactants and is non-spontaneous as written, whereas a negative ΔG° suggests spontaneity and product formation.

What units should ΔG° be in when calculating K, and why?

ΔG° should be in joules per mole (J/mol) because the gas constant R is in J/(mol·K). If ΔG° is given in kJ/mol, it must be converted to J/mol before calculation.

If the reaction's ΔG° were negative instead of positive, how would the equilibrium constant change?

A negative ΔG° would increase the value of K, indicating the reaction favors products and is spontaneous under standard conditions.

Additional Resources

A Reaction Has A Standard Free Energy Change Of 11.60 kJ mol^{-1} (2.772 kcal mol^{-1}). Calculate The Equilibrium — this intriguing problem offers an excellent opportunity to explore the fundamental relationships between thermodynamic parameters and chemical equilibrium. Understanding how the standard free energy change (ΔG°) influences the position of equilibrium is essential in chemistry, especially in fields like physical chemistry, biochemistry, and chemical engineering. This guide will walk you through the concepts, formulas, step-by-step calculations, and practical implications involved in solving such a problem, ensuring you gain a comprehensive understanding of the underlying principles.

Understanding the Key Concepts

Before diving into calculations, it’s important to clarify some core thermodynamic concepts:

- Standard Free Energy Change (ΔG°): The Gibbs free energy change for a reaction under standard conditions (usually 1 bar pressure, 1 M concentration for solutions, and specified temperature). It indicates whether a reaction is spontaneous (ΔG° < 0) or non-spontaneous (ΔG° > 0).

- Equilibrium Constant (K): A quantitative measure of the position of equilibrium for a

reversible reaction. It relates the concentrations or partial pressures of reactants and products at equilibrium.

- Relationship Between ΔG° and K: The fundamental thermodynamic relation connecting free energy change and equilibrium constant is:

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

where:
- R = universal gas constant,
- T = temperature in Kelvin,
- $\ln$ = natural logarithm,
- K = equilibrium constant.

This relation forms the backbone of many calculations aimed at determining the equilibrium state from thermodynamic data.

Step 1: Recognize the Given Data

The problem states:

- Standard free energy change, $\Delta G^\circ = 11.60\, \mathrm{kJ\, mol^{-1}}$
- Alternatively, in calories: $2.772\, \mathrm{kcal\, mol^{-1}}$

Note: Since the calculation involves natural logs and the gas constant, it’s best to work with SI units (kilojoules).

Step 2: Convert Units and Constants

Given:

- $\Delta G^\circ = 11.60\, \mathrm{kJ\, mol^{-1}}$
- $R = 8.314\, \mathrm{J\, mol^{-1}\, K^{-1}}$

To match units:

$$
\Delta G^\circ = 11.60\, \mathrm{kJ\, mol^{-1}} = 11{,}600\, \mathrm{J\, mol^{-1}}
$$

The temperature (T) is not specified explicitly in the problem statement, which suggests that calculations are to be performed generally or at a standard temperature, often 25°C (298.15 K). If the problem specifies otherwise, adjust accordingly.

Step 3: Use the Thermodynamic Relation to Find K

The key formula:

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

Rearranged to solve for K:

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

Plugging in the known values:

$$
K = e^{-\frac{11,600\, \mathrm{J\, mol^{-1}}}{8.314\, \mathrm{J\, mol^{-1}\, K^{-1}} \times 298.15\, \mathrm{K}}}
$$

Calculations step-by-step:

1. Calculate the denominator:

$$
8.314 \times 298.15 \approx 2478.9\, \mathrm{J\, mol^{-1}}
$$

2. Calculate the exponent numerator:

$$
-11,600\, \mathrm{J\, mol^{-1}}
$$

3. Compute the exponent:

$$
-\frac{11,600}{2478.9} \approx -4.678
$$

4. Find K:

$$
K = e^{-4.678} \approx 0.0093
$$

Result: The Equilibrium Constant

$$
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$$
K \approx 0.0093
$$
$$
\]

This value indicates that at standard conditions, the reaction favors the reactants significantly, as $K \ll 1$.

Interpreting the Result

- $K < 1$: The equilibrium lies to the left, favoring reactants.
- $K \gg 1$: The equilibrium favors products.
- $K \approx 1$: The reaction is near equilibrium, with comparable concentrations of reactants and products.

In our case, $K \approx 0.0093$, meaning the reaction is predominantly reactant-favored.

Additional Considerations

Effect of Temperature

If the temperature changes, K will also change, as dictated by the same relationship:

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

and the sign and magnitude of ΔG° can shift with temperature, altering the equilibrium position.

Sign of ΔG°

- Positive ΔG°: Non-spontaneous under standard conditions, $K < 1$, favoring reactants.
- Negative ΔG°: Spontaneous, $K > 1$, favoring products.

In this problem, ΔG° is positive, consistent with our calculated low value of K.

Practical Applications

Understanding how to calculate the equilibrium constant from ΔG° is vital across various domains:

- Chemical manufacturing: To predict yields and optimize reaction conditions.
- Biochemistry: To understand enzyme catalysis and metabolic pathway favorability.

- Environmental science: To evaluate pollutant degradation or formation reactions.
- Material science: To assess stability of compounds and materials.

Summary: How to Approach Similar Problems

1. Identify the given thermodynamic data: ΔG°, temperature.
2. Convert units consistently: Joules, Kelvin.
3. Use the relation: $\Delta G^\circ = -RT \ln K$.
4. Solve for K: Exponentiate the negative ratio.
5. Interpret the magnitude of K: To assess the reaction favorability.
6. Consider temperature effects: Remember that both ΔG° and K are temperature-dependent.

Final Thoughts

The relationship between free energy change and equilibrium constant offers a powerful tool for chemists and scientists to predict reaction behavior under standard conditions. By mastering this calculation, you can better understand reaction spontaneity, optimize conditions, and interpret thermodynamic data across a broad spectrum of chemical contexts.

In conclusion, given a standard free energy change of 11.60 kJ mol^{-1}, the equilibrium constant at 25°C (298.15 K) is approximately 0.0093, indicating a reaction that heavily favors reactants. This example underscores the importance of thermodynamic principles in predicting and controlling chemical processes.

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