Calculate The Standard Free Energy Change At 25 Degrees Celsius Given The Equilibrium Constant Of 1.3

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Understanding how to determine the standard free energy change (ΔG°) from the equilibrium constant (K_{eq}) is fundamental in thermodynamics and chemical kinetics. When a chemical reaction reaches equilibrium, the ratio of product and reactant concentrations is characterized by K_{eq}. Knowing this equilibrium constant at a specific temperature allows us to calculate the standard Gibbs free energy change, which provides insight into the spontaneity and thermodynamic favorability of the reaction. In this article, we will explore the theoretical background, the relevant equations, and the step-by-step process to compute ΔG° at 25°C given $K_{eq} = 1.3$.

Theoretical Background

Understanding the Equilibrium Constant (K_{eq})

The equilibrium constant is a dimensionless number that expresses the ratio of concentrations or activities of products to reactants at equilibrium. Its value indicates whether a reaction favors products or reactants:

- If $K_{eq} > 1$, the reaction favors products.
- If $K_{eq} < 1$, the reaction favors reactants.
- If $K_{eq} = 1$, the system is at equilibrium with equal activities of reactants and products.

Standard Gibbs Free Energy Change ($\Delta

G^\circ\))

The standard Gibbs free energy change is the change in free energy when reactants are converted to products under standard conditions (usually 1 bar pressure, 1 M concentrations, and specified temperature). It indicates whether a process is spontaneous (\(\Delta G^\circ < 0\)) or non-spontaneous (\(\Delta G^\circ > 0\)).

Relationship Between \(\Delta G^\circ\) and \(K_{eq}\)

The Fundamental Equation

The relation connecting \(\Delta G^\circ\) and \(K_{eq}\) is derived from thermodynamic principles and is given by:

\(\Delta G^\circ = -RT \ln K_{eq}\)

where:

- \(R\) = universal gas constant (8.314 J/mol·K)
- \(T\) = absolute temperature in Kelvin (K)
- \(\ln\) = natural logarithm

Significance of the Equation

This equation indicates that if \(K_{eq} > 1\), then \(\ln K_{eq}\) is positive, leading to a negative \(\Delta G^\circ\), meaning the reaction is spontaneous under standard conditions. Conversely, if \(K_{eq} < 1\), \(\Delta G^\circ\) is positive, indicating non-spontaneity.

Calculating \(\Delta G^\circ\) at 25°C for \(K_{eq} = 1.3\)

Step 1: Convert Temperature to Kelvin

The temperature given is 25°C. To use the thermodynamic equation, convert Celsius to Kelvin:

$T = 25 + 273.15 = 298.15\ K$

Step 2: Write Down Known Values

- $K_{eq} = 1.3$
- $T = 298.15\ \text{K}$
- $R = 8.314\ \text{J/mol·K}$

Step 3: Calculate $\ln K_{eq}$

Calculate the natural logarithm of the equilibrium constant:

$\ln 1.3 \approx 0.2624$

Step 4: Apply the Equation to Find ΔG°

Use the relation:

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

Substitute the known values:

$\Delta G^\circ = - (8.314\ \text{J/mol·K}) \times 298.15\ \text{K} \times 0.2624$

Step 5: Perform the Calculation

$\Delta G^\circ = -8.314 \times 298.15 \times 0.2624 \approx -8.314 \times 78.23 \approx -650.4\ \text{J/mol}$

Interpreting the Result

Significance of the Calculated ΔG°

The computed standard free energy change is approximately $-650.4\, \text{J/mol}$. The negative sign indicates that the reaction favors the formation of products at 25°C under standard conditions, and the process is thermodynamically spontaneous in the forward direction.

Implications for Reaction Spontaneity

- Since $\Delta G^\circ < 0$, the reaction tends to proceed spontaneously towards equilibrium.
- The magnitude ($~650\, \text{J/mol}$) reflects a moderate drive towards products under standard conditions.

Additional Considerations

Effect of Temperature

The value of ΔG° is temperature-dependent. At different temperatures, the equilibrium constant may change, which would alter the free energy change accordingly. The relation remains valid as long as the temperature is known.

Standard Conditions and Real-World Systems

While the calculation assumes standard conditions, actual concentrations and pressures may differ, resulting in different free energy changes (ΔG). The reaction quotient Q replaces K_{eq} in non-standard situations, and the actual free energy change can be computed using:

$\Delta G = \Delta G^\circ + RT \ln Q$

Summary

- To calculate the standard free energy change (ΔG°) at 25°C, use the relation $\Delta G^\circ = -RT \ln K_{eq}$.
- Convert temperature to Kelvin, obtain the natural logarithm of the

equilibrium constant, and perform the multiplication.

- A negative ΔG° indicates a spontaneous process under standard conditions.

Conclusion

By applying fundamental thermodynamic principles, we have demonstrated how to calculate the standard free energy change at 25°C given an equilibrium constant of 1.3. This process underscores the importance of thermodynamic relationships in predicting reaction spontaneity and understanding chemical equilibria. Mastery of these calculations is essential for chemists, chemical engineers, and anyone involved in analyzing reaction feasibility and designing chemical processes.

Frequently Asked Questions

How do you calculate the standard free energy change (ΔG°) given the equilibrium constant (K) at 25°C?

Use the equation ΔG° = -RT ln K, where R is the gas constant (8.314 J/mol·K), T is the temperature in Kelvin (298 K at 25°C), and K is the equilibrium constant.

What is the value of R (gas constant) in the ΔG° calculation?

The gas constant R is 8.314 J/mol·K.

How do you convert 25°C to Kelvin for thermodynamic calculations?

Add 273.15 to 25°C, so 25°C equals 298.15 K for calculations.

What is the natural logarithm of the equilibrium constant (K = 1.3)?

ln 1.3 ≈ 0.262.

Calculate the standard free energy change (ΔG°) at 25°C with K = 1.3.

ΔG° = - (8.314 J/mol·K) × 298 K × ln(1.3) ≈ -8.314 × 298 × 0.262 ≈ -649 J/mol.

What does a negative ΔG° indicate about the reaction at standard conditions?

A negative ΔG° indicates that the reaction is spontaneous under standard conditions at 25°C.

How does the magnitude of ΔG° relate to the position of equilibrium?

A more negative ΔG° suggests the equilibrium favors products, while a positive ΔG° favors reactants; a ΔG° close to zero indicates equilibrium is nearly balanced.

Additional Resources

Calculate The Standard Free Energy Change At 25 Degrees Celsius Given The Equilibrium Constant Of 1.3

Understanding the thermodynamics of chemical reactions is fundamental to predicting whether a reaction will proceed spontaneously under certain conditions. One of the key parameters in this domain is the standard Gibbs free energy change (ΔG°), which provides insight into the thermodynamic favorability of a reaction under standard conditions. This article delves into how to calculate ΔG° at 25°C (which is 298.15 K) given an equilibrium constant (K). We will explore the theoretical background, the relevant equations, and practical steps, ensuring a comprehensive understanding of the process.

Fundamentals of Thermodynamics in Chemical Reactions

Before diving into calculations, it’s crucial to grasp the basic principles that connect thermodynamics, equilibrium, and free energy.

Gibbs Free Energy (G)

- Definition: Gibbs free energy (G) is a thermodynamic potential that measures the maximum reversible work obtainable from a system at constant temperature and pressure.
- Expression: $G = H - TS$, where:
- H is enthalpy,
- T is temperature (Kelvin),
- S is entropy.

Change in Gibbs Free Energy (ΔG)

- Represents the energy change during a process.
- When $\Delta G < 0$, the process is spontaneous.
- When $\Delta G > 0$, the process is non-spontaneous.
- When $\Delta G = 0$, the system is at equilibrium.

Standard Gibbs Free Energy Change (ΔG°)

- The change in free energy when reactants and products are in their standard states (usually 1 bar or 1 M concentration).
- Indicates the inherent thermodynamic favorability of a reaction under standard conditions.

Relationship Between Equilibrium Constant and ΔG°

The core relation connecting thermodynamics to reaction equilibrium is:

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

Where:

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

This equation forms the foundation for calculating ΔG° from the equilibrium constant.

Calculating ΔG° at 25°C with K = 1.3

Given:

- $K = 1.3$
- Temperature $T = 25^\circ C = 298.15\,K$

Let's walk through the calculation step by step.

Step 1: Convert Temperature to Kelvin

- $T = 25 + 273.15 = 298.15\,K$

Step 2: Plug Values into the Equation

$$
\Delta G^\circ = - (8.314\, \text{J/(mol·K)}) \times 298.15\,K \times \ln(1.3)
$$

Step 3: Calculate the Natural Logarithm of K

$$
\ln(1.3) \approx 0.262
$$

Step 4: Perform the Multiplication

$$
\Delta G^\circ = -8.314 \times 298.15 \times 0.262
$$

$$
\Delta G^\circ \approx -8.314 \times 298.15 \times 0.262 \approx -8.314 \times 78.07 \approx -648.8\, \text{J/mol}
$$

Result:

$$
\boxed{
\Delta G^\circ \approx -649\, \text{J/mol}
}
$$

This indicates that the standard free energy change at 25°C for a reaction with an equilibrium constant of 1.3 is approximately -649 Joules per mole.

Interpreting the Result

- Negative ΔG°: The reaction is thermodynamically favorable under standard conditions, tending toward product formation.
- Magnitude: A ΔG° of around -649 J/mol suggests a reaction that is slightly favorable but not strongly so; reactions with large negative ΔG° values (e.g., -50 kJ/mol) are much more spontaneous.
- Equilibrium Implication: Since $K = 1.3$ is close to 1, the reaction equilibrium lies near the middle, favoring neither reactants nor products strongly.

Additional Considerations and Applications

Understanding how to calculate ΔG° from K at a specific temperature has broad applications:

1. Predicting Reaction Spontaneity

- Using the calculated ΔG°, chemists can determine whether a reaction will proceed spontaneously under standard conditions.
- For reactions with $K > 1$, ΔG° is negative; for $K < 1$, ΔG° is positive.

2. Thermodynamic Data in Reaction Design

- When designing chemical syntheses, knowing ΔG° helps in optimizing conditions to favor desired products.
- Adjustments in temperature can shift the equilibrium; thus, understanding

the temperature dependence of ΔG° is vital.

3. Relationship with Reaction Quotient (Q)

- When the reaction is not at equilibrium, the actual free energy change (ΔG) can be calculated using:

$$
\Delta G = \Delta G^\circ + RT \ln Q
$$

where Q is the reaction quotient. This allows for real-time analysis of reaction progress.

4. Linking to Other Thermodynamic Parameters

- ΔG° is related to enthalpy (ΔH°) and entropy (ΔS°) via:

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

- Knowing ΔG° and K, alongside temperature, can help determine these other thermodynamic parameters.

Limitations and Assumptions in the Calculation

While the calculation presented is straightforward, several assumptions underpin it:

- Standard Conditions: The calculation assumes standard states (1 bar, 1 M concentrations). Deviations can alter ΔG.
- Ideal Behavior: The reaction system behaves ideally; activity coefficients are assumed to be unity.
- Temperature Constancy: The calculation presumes temperature stability at 25°C. Variations can significantly impact ΔG.
- Equilibrium State: The reaction is at equilibrium, and K accurately reflects the system's thermodynamics.

Practical Examples and Broader Context

To contextualize the calculation, consider real-world reactions where such data is vital:

- Biochemical Reactions: Enzyme catalysis often involves reactions with equilibrium constants near 1. Calculating ΔG° helps assess reaction feasibility.
- Industrial Processes: Synthesis of chemicals, pharmaceuticals, or materials relies on thermodynamic insights to optimize yields.
- Environmental Chemistry: Understanding pollutant degradation or resource recovery involves thermodynamic calculations.

Summary and Final Remarks

Calculating the standard free energy change at 25°C given the equilibrium constant is a fundamental skill in thermodynamics and chemistry. The key formula:

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

enables quick and accurate determination of thermodynamic favorability, guiding experimental and theoretical work.

For the specific case where $K = 1.3$, the calculation shows a slightly negative ΔG°, indicating a reaction that is mildly spontaneous under standard conditions at room temperature. Recognizing how small changes in K impact ΔG° emphasizes the sensitivity of reaction spontaneity to equilibrium constants.

This understanding informs not only academic studies but also practical applications in chemical manufacturing, biochemical pathways, and environmental science.

In conclusion, mastering the calculation of ΔG° from K at specific temperatures like 25°C is essential for predicting reaction behavior, designing experiments, and interpreting thermodynamic data across various fields of chemistry and related sciences.

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