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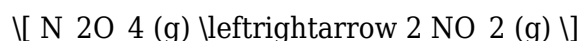
For The Reaction $\text{N}_2\text{O}_4 = 2\text{NO}_2$, $K_p = 0.148$ At A Temperature Of 298K. What Is The K_p For The Following

Understanding chemical equilibrium is fundamental in chemistry, especially when it comes to predicting how reactions behave under different conditions. In this article, we explore the equilibrium constant (K_p) for the reaction involving nitrogen tetroxide (N_2O_4) and nitrogen dioxide (NO_2) at 298 K. We will analyze how to determine (K_p) for various related reactions, interpret the significance of equilibrium constants, and provide practical examples for students and professionals alike. Whether you're studying for exams or working on industrial processes, grasping these concepts is essential for a deep understanding of chemical equilibria.

Understanding the Reaction and Its Equilibrium Constant (K_p)

The Reaction Under Consideration

The fundamental reaction we are examining is:



This reaction is reversible and reaches an equilibrium point where the rates of the forward and reverse reactions are equal.

Given Data

- Equilibrium constant in terms of partial pressures: $(K_p = 0.148)$
- Temperature: 298 K

The (K_p) value indicates the relative concentrations (partial pressures) of (N_2O_4) and (NO_2) at equilibrium under the specified temperature.

Significance of the Equilibrium Constant (K_p)

What Does K_p Represent?

- K_p is the ratio of the partial pressures of products to reactants, each raised to the power of their coefficients in the balanced equation.
- For the reaction above:

$$K_p = \frac{(P_{\text{NO}_2})^2}{P_{\text{N}_2\text{O}_4}}$$

Interpreting the Value $K_p = 0.148$

- A K_p less than 1 indicates that, at equilibrium, the reactants (N_2O_4) are favored over the products (NO_2).
- Conversely, a K_p greater than 1 suggests a product-favored reaction.

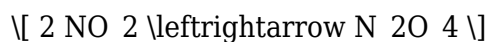
Understanding this helps predict how shifts in conditions might affect the composition of the equilibrium mixture.

Calculating K_p for Related Reactions

Often, chemists need to determine the equilibrium constant for reactions related to the original one, especially when dealing with complex systems or multiple steps. Below are common scenarios and how to approach them.

1. Reversing the Reaction

Suppose the reaction is written in reverse:



- The equilibrium constant for this reverse reaction, $K_{p_{\text{rev}}}$, is the reciprocal of the original:

$$K_{p_{\text{rev}}} = \frac{1}{K_p} = \frac{1}{0.148} \approx 6.76$$

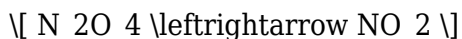
- Key Point: Reversing a reaction inverts its K_p .

2. Adjusting for Stoichiometric Changes

Consider a reaction where the coefficients are scaled:



Suppose we are asked about:



- Since the original reaction produces 2 moles of NO_2 , the equilibrium constant for producing only 1 mole of NO_2 (by taking the square root) is:

$$K_p^{(1)} = \sqrt{K_p} = \sqrt{0.148} \approx 0.385$$

- Key Point: When the coefficients are scaled, the new K_p is obtained by raising the original K_p to the power corresponding to the change in stoichiometry.

3. Combining Reactions

Suppose two reactions are combined:



which is the reverse of Reaction 1 with coefficients scaled appropriately.

- The overall K_p for combined reactions is the product of their individual K_p values, considering reversals.

Practical Examples: Calculating K_p for Variations

Let's explore some common scenarios and how to compute the equilibrium constants.

Example 1: Determining K_p for the Reaction $\text{N}_2\text{O}_4 \rightleftharpoons 2 \text{NO}_2$ at Different Conditions

Given the original $K_p = 0.148$ at 298 K, what if the temperature changes? Since K_p depends on temperature, the Van 't Hoff equation can estimate the change.

$$\ln \frac{K_{p2}}{K_{p1}} = - \frac{\Delta H^\circ}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

- Here, ΔH° is the enthalpy change for the reaction, R is the gas constant, and T_1 ,

T_1 and T_2 are the initial and final temperatures.

Note: Without ΔH° , precise calculation isn't possible, but understanding the relation helps in thermodynamic predictions.

Example 2: Calculating Partial Pressures at Equilibrium

Suppose the total pressure in a reaction vessel is 1 atm, and the reaction mixture is at equilibrium with partial pressures $P_{\text{N}_2\text{O}_4}$ and P_{NO_2} .

Given:

$$K_p = \frac{(P_{\text{NO}_2})^2}{P_{\text{N}_2\text{O}_4}} = 0.148$$

$$P_{\text{total}} = P_{\text{N}_2\text{O}_4} + P_{\text{NO}_2}$$

Assuming x is the partial pressure of NO_2 :

$$P_{\text{NO}_2} = x$$

$$P_{\text{N}_2\text{O}_4} = y$$

From the relation:

$$K_p = \frac{x^2}{y}$$

And:

$$x + y = 1 \text{ atm}$$

Express $y = 1 - x$, then:

$$0.148 = \frac{x^2}{1 - x}$$

Solve for x to find the partial pressures at equilibrium.

Implications for Industrial and Laboratory Applications

Understanding how to calculate and manipulate (K_p) values is vital in various applications:

- Industrial Synthesis: Controlling the production of nitrogen dioxide and tetroxide based on desired equilibrium compositions.
- Environmental Chemistry: Modeling nitrogen oxides in the atmosphere, which impact air quality.
- Laboratory Experiments: Designing experiments that require specific concentrations of reactive gases.

Conclusion: Mastering (K_p) Calculations and Applications

In summary, the equilibrium constant (K_p) provides crucial insights into the behavior of reversible reactions like $(N_2O_4 \rightleftharpoons 2 NO_2)$. Starting from a known (K_p) value at a specific temperature, chemists can determine the constants for related reactions, reverse reactions, and different conditions using mathematical principles such as reciprocals, stoichiometric adjustments, and thermodynamic equations. Mastery of these concepts enables accurate prediction of reaction behavior, optimization of industrial processes, and a deeper understanding of chemical equilibria. Whether you're analyzing environmental systems or designing chemical syntheses, knowing how to manipulate and interpret (K_p) values is an indispensable skill in modern chemistry.

Key Points to Remember:

- Reversing a reaction inverts (K_p) .
- Scaling coefficients raises (K_p) to the corresponding power.
- Combining reactions involves multiplying their (K_p) values.
- Temperature changes affect (K_p) , often estimated via the Van 't Hoff equation.
- Practical calculations often involve setting up equations based on partial pressures and total pressure.

By understanding these principles, you will be better equipped to analyze and predict the outcomes of chemical reactions involving nitrogen oxides and similar systems.

SEO Keywords: (K_p) calculation, nitrogen tetroxide, nitrogen dioxide, chemical equilibrium, reaction quotient, equilibrium constant, thermodynamics, reaction analysis, industrial chemistry, environmental chemistry

Frequently Asked Questions

What is the equilibrium constant K_p for the reaction $\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$ at 298K?

The given K_p for the reaction at 298K is 0.148.

How does temperature affect the value of K_p for the reaction $\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$?

Since the temperature is provided as 298K, the K_p value of 0.148 applies at this temperature; changing the temperature would alter the K_p depending on the reaction's enthalpy change.

What is the significance of the K_p value in the reaction $\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$?

K_p indicates the ratio of partial pressures of products to reactants at equilibrium; a K_p of 0.148 suggests the reactant N_2O_4 is favored at 298K.

How would you calculate the equilibrium partial pressures of N_2O_4 and NO_2 given initial conditions?

You would set up an equilibrium expression using $K_p = (P_{\text{NO}_2})^2 / P_{\text{N}_2\text{O}_4}$, and use initial pressures and the change in moles to solve for the equilibrium pressures.

If the reaction shifts to produce more NO_2 , how would the K_p value change at 298K?

The K_p value is temperature-dependent; at constant temperature of 298K, it remains unchanged. If the temperature stays the same, K_p remains at 0.148 regardless of the shift.

Can you determine the K_p for a different reaction involving N_2O_4 or NO_2 from the given data?

No, the given K_p of 0.148 specifically pertains to the reaction $\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$ at 298K; different reactions or conditions require separate equilibrium calculations.

Additional Resources

Understanding the Equilibrium: $\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$ and Its K_p at 298K

When exploring chemical equilibria, the equilibrium constant (K_p) provides crucial insight into the relative amounts of reactants and products at a given temperature. In particular, the reaction $\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$ serves as a classic example for understanding how temperature influences equilibrium and

how to calculate the value of K_p under different conditions. With a known K_p of 0.148 at 298 K, we can delve into the principles governing this reaction, interpret what the equilibrium constant tells us, and explore how to determine K_p under altered conditions or for related reactions.

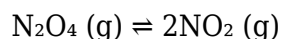
This article aims to provide a comprehensive guide to analyzing this reaction, emphasizing the importance of equilibrium constants, the factors influencing them, and detailed steps for calculating K_p for various scenarios. Whether you're a chemistry student, researcher, or enthusiast, this detailed breakdown will enhance your understanding of chemical equilibria and their practical applications.

Understanding the Reaction: $\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$

Before diving into calculations, it's essential to comprehend the reaction itself:

- N_2O_4 (dinitrogen tetroxide) is a colorless, relatively stable gas at room temperature.
- NO_2 (nitrogen dioxide) is a reddish-brown gas known for its toxic nature and distinct coloration.
- The reaction is reversible and reaches a dynamic equilibrium in a closed system, where the rates of the forward and reverse reactions are equal.

The reaction can be represented as:



In the context of equilibrium, the equilibrium constant K_p relates the partial pressures of gases at equilibrium:

$$K_p = \frac{(P_{\text{NO}_2})^2}{P_{\text{N}_2\text{O}_4}}$$

where:

- P_{NO_2} is the partial pressure of NO_2 ,
- $P_{\text{N}_2\text{O}_4}$ is the partial pressure of N_2O_4 .

Significance of the Known K_p at 298K

Given:

- $K_p = 0.148$ at 298 K

This value indicates that at room temperature, the equilibrium favors the reactant N_2O_4 , since the equilibrium constant is less than 1. A small K_p suggests that most of the nitrogen dioxide has recombined to form N_2O_4 at this temperature.

Understanding how this equilibrium constant changes with temperature or under different conditions is vital for industrial processes, safety considerations, and fundamental chemistry.

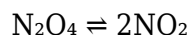
Factors Affecting the Equilibrium Constant (K_p)

The value of K_p depends primarily on temperature. According to Le Châtelier's principle and thermodynamics:

- Endothermic reaction (absorbs heat): Increasing temperature increases K_p.
- Exothermic reaction (releases heat): Increasing temperature decreases K_p.

Determining the Enthalpy Change (ΔH°)

For the reaction:



The enthalpy change (ΔH°) can be found in thermodynamic tables or literature. It is approximately:

- ΔH° ≈ +57 kJ/mol (positive, indicating endothermic)

This positive ΔH° suggests that increasing temperature will favor the formation of NO₂, increasing K_p.

Calculating K_p at Different Temperatures

Suppose you need to find the K_p at a different temperature, say 350 K. The relationship between K_p and temperature is given by the van 't Hoff equation:

$$\ln \left(\frac{K_{p_2}}{K_{p_1}} \right) = -\frac{\Delta H^\circ}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

Where:

- K_{p_1} and K_{p_2} are the equilibrium constants at temperatures T_1 and T_2 ,
- R is the gas constant (8.314 J/(mol·K)),
- ΔH° is in Joules per mol.

Step-by-step Calculation:

1. Identify known values:

- $K_{p_1} = 0.148$ at $T_1 = 298\text{ K}$,
- $T_2 = 350\text{ K}$,
- $\Delta H^\circ = 57,000\text{ J/mol}$,
- $R = 8.314\text{ J/(mol·K)}$.

2. Calculate the change in reciprocal temperature:

$$\frac{1}{T_2} - \frac{1}{T_1} = \frac{1}{350} - \frac{1}{298} \approx 0.002857 - 0.003356 = -0.000499\text{ K}^{-1}$$

\]

3. Apply the van 't Hoff equation:

\[

$$\ln \left(\frac{K_{p_2}}{0.148} \right) = -\frac{57,000}{8.314} \times (-0.000499)$$

\]

\[

$$\ln \left(\frac{K_{p_2}}{0.148} \right) \approx -6844 \times (-0.000499) \approx 3.42$$

\]

4. Solve for (K_{p_2}) :

\[

$$\frac{K_{p_2}}{0.148} = e^{3.42} \approx 30.6$$

\]

\[

$$K_{p_2} = 0.148 \times 30.6 \approx 4.53$$

\]

Result:

At 350 K, the K_p increases to approximately 4.53, indicating a shift toward more NO_2 at higher temperature.

Interpreting the Results: How Temperature Influences Equilibrium

The significant increase in K_p with temperature highlights the endothermic nature of the dissociation:

- At 298 K: $K_p = 0.148$, favoring N_2O_4 .
- At 350 K: $K_p \approx 4.53$, favoring NO_2 .

This shift is consistent with Le Châtelier's principle: adding heat shifts equilibrium toward the endothermic side (producing more NO_2).

Practical Applications and Considerations

Understanding how to manipulate K_p is vital in various contexts:

Industrial Synthesis

- N_2O_4 and NO_2 are involved in nitric acid production. Precise control of temperature ensures optimal yields.
- Adjusting temperature can influence the ratio of N_2O_4 to NO_2 , affecting process efficiency.

Safety Implications

- NO₂ is toxic and reddish-brown, making it easier to detect visually.
- Higher temperatures lead to more NO₂, which can pose safety hazards in storage and handling.

Environmental Impact

- The equilibrium shifts can influence emission levels in combustion or industrial processes.

Summary of Key Points

- The reaction $\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$ is dynamic and temperature-dependent.
- The known K_p at 298 K is 0.148, favoring N_2O_4 .
- Thermodynamic principles, via the van 't Hoff equation, allow estimation of K_p at other temperatures.
- Increasing temperature shifts the equilibrium toward NO₂ in this endothermic reaction.
- Practical control of temperature impacts product yields, safety, and environmental factors.

Additional Related Calculations and Scenarios

1. Calculating K_p at 273 K

Using the same approach with $(T_1 = 298\text{ K})$ and $(T_2 = 273\text{ K})$:

$$\frac{1}{273} - \frac{1}{298} \approx 0.003661 - 0.003356 = 0.000305\text{ K}^{-1}$$

$$\ln \left(\frac{K_{p,273}}{0.148} \right) = -6844 \times 0.000305 \approx -2.09$$

$$K_{p,273} = 0.148 \times e^{-2.09} \approx 0.148 \times 0.124 \approx 0.0183$$

Conclusion:

At 273 K, K_p drops significantly, favoring N_2O_4 even more.

2. Determining the Partial Pressures at Equilibrium

Suppose at 298 K, the partial pressure of N_2O_4 is measured at 0.5 atm. To find the partial pressure of NO₂:

$$K_p = \frac{(P_{\text{NO}_2})^2}{P_{\text{N}_2\text{O}_4}} \rightarrow P_{\text{NO}_2} = \sqrt{K_p \times P_{\text{N}_2\text{O}_4}}$$

$$P_{\text{NO}_2} = \sqrt{0.148 \times 0.5} \approx \sqrt{0.074} \approx 0.272 \text{ atm}$$

This calculation helps in predicting gas concentrations in practical settings.

Final Thoughts: Mastering Equilibrium Calcul

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