CheggSolutions - Thegdp

Chemistry

Topic: Chemical Equilibrium

Given:

$K = 1.4 \times 10^{-9}$
Initial $[A] = 0.24 \text{ mol/L}$
Initial $[B] = 0.36 \text{ mol/L}$

For the equilibrium reaction:
$3A(g) + B(g) \rightleftharpoons C(g) + 2D(g)$

To find: Concentrations of all chemicals at equilibrium.

Step-by-Step Solution:

Step 1: Write down the equilibrium expression

The equilibrium constant expression K for the reaction:
$K = \frac{[C][D]^2}{[A]^3[B]}$

Explanation:

The equilibrium constant expression is derived based on the balanced chemical equation, mixing the products over reactants, each raised to the power of their coefficients in the balanced equation.

Step 2: Set up the ICE Table

	[A] (mol/L)	[B] (mol/L)	[C] (mol/L)	[D] (mol/L)
I	0.24	0.36	0	0
C	-3x	-x	+x	+2x
E	0.24 - 3x	0.36 - x	x	2x

Explanation:

The ICE table setup helps to visualize the change during equilibrium from the initial concentration to the equilibrium concentration by introducing a variable x.

Step 3: Substitute the equilibrium concentrations into the equilibrium expression

Substitute the equilibrium concentrations into the expression $[C] = x$, $[D] = 2x$, $[A] = 0.24 - 3x$ and $[B] = 0.36 - x$:

$K = \frac{[C][D]^2}{[A]^3[B]}$

$K = \frac{x \cdot (2x)^2}{(0.24 - 3x)^3 (0.36 - x)}$

Explanation:

To solve for x, assuming that x is very small (due to the small value of K), simplifies the calculation.

Step 4: Simplify and solve for x

Assume x is very small, thus $0.24 - 3x \approx 0.24$ and $0.36 - x \approx 0.36$:

$1.4 \times 10^{-9} = \frac{4x^3}{0.0248832}$

$1.4 \times 10^{-9} * 0.0248832 = 4x^3$

$3.483648 \times 10^{-11} = 4x^3$

$x^3 = \frac{3.483648 \times 10^{-11}}{4}$

$x^3 = 8.70912 \times 10^{-12}$

$x = \sqrt[3]{8.70912 \times 10^{-12}}$

$x = 2.058 \times 10^{-4} \text{ mol/L}$

Explanation:

By simplifying and solving the cubic equation, the equilibrium concentration for [C] (which is x) is calculated.

Step 5: Calculate the equilibrium concentrations of all species

Using $x = 2.058 \times 10^{-4}$:

- $[A]_{eq} = 0.24 - 3(2.058 \times 10^{-4}) = 0.24 - 6.174 \times 10^{-4} = 0.239382 \text{ mol/L}$
- $[B]_{eq} = 0.36 - 2.058 \times 10^{-4} = 0.359794 \text{ mol/L}$
- $[C]_{eq} = 2.058 \times 10^{-4} \text{ mol/L}$
- $[D]_{eq} = 2(2.058 \times 10^{-4}) = 4.116 \text{times} 10^{-4} \text{ mol/L}$

Explanation:

By substituting x back into the equilibrium concentrations from the ICE table, the final concentrations for all species are calculated.

Final Solution

$[A]_{eq} = 0.239382 \text{ mol/L}$
$[B]_{eq} = 0.359794 \text{ mol/L}$
$[C]_{eq} = 2.058 \times 10^{-4} \text{ mol/L}$
$[D]_{eq} = 4.116 \times 10^{-4} \text{ mol/L}$