

Chemistry

Topic: Chemical Equilibrium

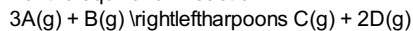
Given:

$$K = 1.4 \times 10^{-9}$$

$$\text{Initial [A]} = 0.24 \text{ mol/L}$$

$$\text{Initial [B]} = 0.36 \text{ mol/L}$$

For the equilibrium reaction:



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:

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

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 $[\text{C}] = x$, $[\text{D}] = 2x$, $[\text{A}] = 0.24 - 3x$ and $[\text{B}] = 0.36 - x$:

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

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

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 $0.36 - x \approx 0.36$:

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

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

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

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

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

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

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

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

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

Explanation:

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

Final Solution

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