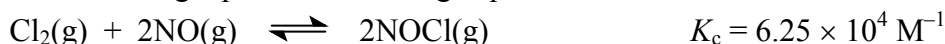


Marks
6

- Consider the following equilibrium in the gas-phase at 35 °C.



Equimolar amounts of NOCl(g) and Cl₂(g) are introduced into a sealed 1.00 L flask. When the system reaches equilibrium at 35 °C, the concentration of NO(g) in the flask is 4.04×10^{-4} M. What amount of Cl₂(g) (in mol) was initially added to the flask?

If the initial concentrations of Cl₂(g) and NOCl (g) are both equal to x , the reaction table is:

	Cl ₂ (g)	2NO(g)	\rightleftharpoons	2NOCl(g)
initial	x	0		x
change	$+y$	$+2y$		$-2y$
equilibrium	$x + y$	$2y$		$x - 2y$

(As [NO(g)] has increased so must [Cl₂(g)] but by half as much since every two NO molecules are made for every Cl₂. [NOCl(g)] must have decreased by the same amount as [NO(g)] has increased. The reaction is proceeding *backwards* from the point at which it has begun.)

As $[\text{NO}(\text{g})]_{\text{equilibrium}} = 2y = 4.04 \times 10^{-4} \text{ M}$, $y = 2.02 \times 10^{-4} \text{ M}$ and so:

$$[\text{Cl}_2(\text{g})]_{\text{equilibrium}} = x + y = x + 2.02 \times 10^{-4} \text{ M} \text{ and}$$

$$[\text{NOCl}(\text{g})]_{\text{equilibrium}} = x - 2y = x - 4.04 \times 10^{-4} \text{ M}$$

From the equilibrium constant:

$$K_c = \frac{[\text{NOCl}(\text{g})]^2}{[\text{Cl}_2(\text{g})][\text{NO}(\text{g})]^2} = \frac{(x - 4.04 \times 10^{-4})^2}{(x + 2.02 \times 10^{-4})(4.04 \times 10^{-4})^2}$$

As the equilibrium for the *forward* reaction is very large (6.25×10^4), the amount of Cl₂ and NO that are produced in the *backward* reaction is very small. It can be assumed that this amount (4.04×10^{-4}) is small compared to x .

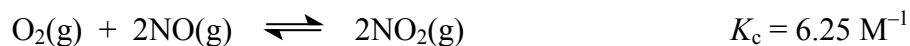
Hence, $x + 2.02 \times 10^{-4} \sim x$ and $x - 4.04 \times 10^{-4} \sim x$:

$$K_c = \frac{(x - 4.04 \times 10^{-4})^2}{(x + 2.02 \times 10^{-4})(4.04 \times 10^{-4})^2} \sim \frac{x^2}{x(4.04 \times 10^{-4})^2} = \frac{x}{(4.04 \times 10^{-4})^2}$$

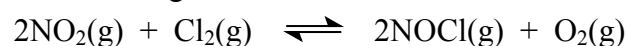
As $K_c = 6.25 \times 10^4$, $[\text{Cl}_2(\text{g})]_{\text{initial}} = x = 0.0102 \text{ M}$. As the volume of the flask is 1.00 L, the number of moles = concentration \times volume = $0.0102 \times 1.00 = 0.0102 \text{ mol}$

Answer: **0.0102 M**

At the same temperature (35 °C) $\text{O}_2(\text{g})$ reacts with $\text{NO}(\text{g})$ according to the equation:



Determine K_c for the following reaction.



The equilibrium constants for the three reactions are:

$$K_c(1) = \frac{[\text{NOCl}]^2}{[\text{Cl}_2][\text{NO}]^2}, \quad K_c(2) = \frac{[\text{NO}_2]^2}{[\text{O}_2][\text{NO}]^2} \quad \text{and} \quad K_c(3) = \frac{[\text{NOCl}]^2[\text{O}_2]}{[\text{NO}_2]^2[\text{Cl}_2]}$$

so that

$$\frac{K_c(1)}{K_c(2)} = \frac{[\text{NOCl}]^2}{[\text{Cl}_2][\text{NO}]^2} \cdot \frac{[\text{NO}_2]^2}{[\text{O}_2][\text{NO}]^2} = \frac{[\text{NOCl}]^2[\text{O}_2]}{[\text{NO}_2]^2[\text{Cl}_2]} = K_c(3)$$

$$\text{Hence, } K_c(3) = \frac{6.25 \times 10^4}{6.25} = 1.00 \times 10^4.$$

$$K_c = 1.00 \times 10^4 \text{ M}$$