Marks 6

- Consider the following equilibrium in the gas-phase at 35 ºC.

$$Cl_2(g) + 2NO(g) \rightleftharpoons 2NOCl(g) \qquad K_c = 6.25 \times 10^4 \text{ M}^{-1}$$

Equimolar amounts of NOCl(g) and Cl_2(g) are introduced into a sealed 1.00 L flask. When the system reaches equilibrium at 35 ºC, the concentration of NO(g) in the flask is 4.04×10^{-4} M. What amount of Cl_2(g) (in mol) was initially added to the flask?

If the initial concentrations of Cl_2(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 [Cl_2(g)] but by half as much since every two NO molecules are made for every Cl_2. [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 $[NO(g)]_{equilibrium} = 2y = 4.04 \times 10^{-4}$ M, $y = 2.02 \times 10^{-4}$ M and so:

$[Cl_2(g)]_{equilibrium} = x + y = x + 2.02 \times 10^{-4}$ M and
$[NOCl(g)]_{equilibrium} = x - 2y = x - 4.04 \times 10^{-4}$ M

From the equilibrium constant:

$$K_c = \frac{[NOCl(g)]^2}{[Cl_2(g)][NO(g)]^2} = \frac{(x - 4.04\text{x}10^{-4})^2}{(x + 2.02\text{x}10^{-4})(4.04\text{x}10^{-4})^2}$$

As the equilibrium for the *forward* reaction is very large (6.25 x 10^4), the amount of Cl_2 and NO that are produced in the *backward* reaction is very small. It can be assumed that this amount (4.04 x 10^{-4}) is small compared to x.

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

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

As $K_c = 6.25 \times 10^4$, $[Cl_2(g)]_{initial} = x = 0.0102$ M. As the volume of the flask is 1.00 L, the number of moles = concentration × volume = 0.0102 × 1.00 = 0.0102 mol

Answer: **0.0102 M**

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At the same temperature (35 ºC) $O_2(g)$ reacts with NO(g) according to the equation:

$$O_2(g) + 2NO(g) \rightleftharpoons 2NO_2(g) \qquad K_c = 6.25\ M^{-1}$$

Determine K_c for the following reaction.

$$2NO_2(g) + Cl_2(g) \rightleftharpoons 2NOCl(g) + O_2(g)$$

The equilibrium constants for the three reactions are:

$$K_c\,(1) = \frac{[NOCl]^2}{[Cl_2][NO]^2}\ ,\ K_c\,(2) = \frac{[NO_2]^2}{[O_2][NO]^2}\ \text{ and }\ K_c\,(3) = \frac{[NOCl]^2[O_2]}{[NO_2]^2[Cl_2]}$$

so that

$$\frac{K_c\,(1)}{K_c\,(2)} = \frac{[NOCl]^2}{[Cl_2][NO]^2} \Big/ \frac{[NO_2]^2}{[O_2][NO]^2} = \frac{[NOCl]^2[O_2]}{[NO_2]^2[Cl_2]} = K_c(3)$$

Hence, $K_c\,(3) = \dfrac{6.25\times10^{4}}{6.25} = 1.00 \times 10^{4}$.

K_c = **1.00 × 10**$^{-4}$ **M**

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