Marks
5

If 0.250 mol of HI(g) is introduced into a 2.00 L flask at 700 °C, what will be the concentration of I_2(g) at equilibrium?

The initial concentration of HI(g) is 0.250 / 2.00 mol L^{-1} = 0.125 mol L^{-1}.

	H_2(g)	I_2(g)	$\rightleftharpoons$	2HI(g)
Initial	**0**	**0**		**0.125**
Change	**+x**	**+x**		**-2x**
Equilibrium	**x**	**x**		**0.125 - 2x**

Thus,

$$K_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{(0.125-2x)^2}{(x)(x)} = \frac{(0.125-2x)^2}{x^2} = 49.0 \text{ (from 2008-N-5)}$$

$$(49.0)^{1/2} = \frac{(0.125-2x)}{x}$$

Rearranging gives x = [I_2(g)] = 0.0139 M.

Answer: **0.0139 M**

If 0.274 g of H_2S were now introduced into the same flask, what would be the concentration of S_2(g) at equilibrium?

The molar mass of H_2S is (2 × 1.008 (H) + 32.06 (S)) = 34.08 g mol^{-1}. Hence, 0.274 g of H_2S corresponds to:

number of moles = mass / molar mass
= (0.274 g) / (34.08 g mol^{-1}) = 8.04 × 10^{-3} mol

The initial concentration of H_2S is thus 8.04 × 10^{-3} mol / 2.00 M = 4.02 × 10^{-3} M.

From above, [I_2(g)] = 0.0139 M and [HI(g)] = (0.125 – 2 × 0.0139) M = 0.0972 M.

Using the overall equilibrium reaction derived in 2008-N-5:

	2I_2(g)	2H_2S(g)	$\rightleftharpoons$	S_2(g)	4HI(g)
Initial	**0.0139**	**0.00402**		**0**	**0.0972**
Change	**-2x**	**-2x**		**+x**	**+4x**
Equilibrium	**0.0139 - 2x**	**0.00402 – 2x**		**x**	**0.0972 + 4x**

ANSWER CONTINUES ON THE NEXT PAGE

Thus,

$$K_c = \frac{[\mathbf{S_2}][\mathbf{HI}]^4}{[\mathbf{I_2}]^2[\mathbf{H_2S}]^2} = \frac{(x)(0.0972+4x)^4}{(0.0139-2x)^2(0.00402-2x)^2}$$

$$\sim \frac{(x)(0.0972)^4}{(0.0139)^2(0.00402)^2} = 2.23 \times 10^{-5} \quad \textbf{(from 2008-N-5)}$$

where the small x approximation has been used as K_c is so small. This gives:

$x = [S_2(g)] = 7.82 \times 10^{-10}$ **M**

Answer: $\mathbf{7.82 \times 10^{-10}}$ **M**