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)	⇌	2HI(g)
Initial	**0**	**0**		**0.125**
Change	**+*x***	**+*x***		**-2*x***
Equilibrium	***x***	***x***		**0.125 - 2*x***

Thus,

$$K_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{(0.125-2x)^2}{(x)(x)} = \frac{(0.125-2x)^2}{x^2} = 49.0$$ **(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 = 0.00972 M.

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

	$2I_2$(g)	$2H_2S$(g)	⇌	S_2(g)	4HI(g)
Initial	**0.0139**	**0.00402**		**0**	**0.00972**
Change	**-2*x***	**-2*x***		**+*x***	**+2*x***
Equilibrium	**0.0139 - 2*x***	**0.00402 – 2*x***		***x***	**0.00972 + 4*x***

ANSWER CONTINUES ON THE NEXT PAGE

Thus,

$$K_c = \frac{[S_2][HI]^4}{[I_2]^2[I_2]^2} = \frac{(x)(0.00972+4x)^4}{(0.0139-2x)^2(0.00402-2x)^2}$$
$$\sim \frac{(x)(0.00972)^4}{(0.0139)^2(0.00402)^2} = 2.23 \times 10^{-5} \text{ (from 2008-N-5)}$$

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: 7.82×10^{-10} **M**