Marks
4

- At 700 °C, hydrogen and iodine react according to the following equation.

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g) \qquad K_c = 49.0$$

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,

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

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

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

Answer: **0.0139 M**

Hydrogen also reacts with sulfur at 700 °C:

$$2H_2(g) + S_2(g) \rightleftharpoons 2H_2S(g) \qquad K_c = 1.075 \times 10^8$$

Determine K_c for the following overall equilibrium reaction at 700 °C.

$$2I_2(g) + 2H_2S(g) \rightleftharpoons S_2(g) + 4HI(g)$$

The overall reaction corresponds to the twice the first reaction combined with the reverse of the second reaction:

$2H_2(g) + 2I_2(g) \rightleftharpoons 4HI(g)$	**$K_c(1) = (49.0)^2$**
$2H_2S(g) \rightleftharpoons 2H_2(g) + S_2(g)$	**$K_c(2) = 1/(1.075 \times 10^8)$**
$2I_2(g) + 2H_2S(g) \rightleftharpoons S_2(g) + 4HI(g)$	**$K_c(3) = K_c(1) \times K_c(2)$**

The 1st reaction is doubled so the original equilibrium constant is squared.

The 2nd reaction is reversed so the reciprocal of the equilibrium constant is used.

ANSWER CONTINUES ON THE NEXT PAGE

The two reactions are then combined and the overall equilibrium constant is then the product:

$$\boldsymbol{K_c(3) = K_c(1) \times K_c(2) = (49.0)^2 \times (1/(1.075 \times 10^8) = 2.23 \times 10^{-5}}$$

Answer: **2.23×10^{-5}**