

15. The equilibrium constant for the reaction:



at 448°C is 50.

a) If 1.0 mol of H_2 is mixed with 1.0 mol of I_2 in a 0.50 L container and allowed to react at 448°C, what is the equilibrium $[\text{HI}]$?

Initial $[\text{H}_2] = \frac{1.0 \text{ mol}}{0.50 \text{ L}} = 2.0 \text{ M}$, $[\text{I}_2] = 2.0 \text{ M}$, $[\text{HI}] = 0$



[I]	2.0	2.0	0
[C]	-x	-x	+2x
[E]	2.0-x	2.0-x	2x
[E]	2.0-1.559	2.0-1.559	2(1.559)
[E]	0.44	0.44	3.1

$$K_{eq} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$$

$$50. = \frac{(2x)^2}{(2.0-x)^2}$$

$$\sqrt{50.} = \frac{2x}{2-x}$$

$$\begin{aligned} 7.071(2-x) &= 2x \\ 14.142 - 7.071x &= 2x \\ 14.142 &= 9.071x \\ x &= \frac{14.142}{9.071} = 1.559 \end{aligned}$$

Answer $[\text{HI}] = 3.1 \text{ M}$

b) How many moles of HI are formed at equilibrium? (Actual yield)

$$3.118 \text{ M} \times 0.50 \text{ L} = 1.6 \text{ mol}$$

Answer 1.6 moles of HI

16. Given K_{eq} for the reaction:is 0.042 at 250°C, what will happen if 2.50 mol of PCl_5 , 0.600 mol of Cl_2 and 0.600 mol of PCl_3 are placed in a 1.00 flask at 250°C? (Will the reaction shift left, right, or not occur at all?)

$$K_{eq} = \frac{[\text{PCl}_3][\text{Cl}_2]}{[\text{PCl}_5]}$$

$$\text{Trial } K_{eq} = \frac{(0.600)(0.600)}{(2.50)}$$

$$\text{Trial } K_{eq} = 0.144$$

$$\text{Actual } K_{eq} = 0.042$$

Since Trial $K_{eq} > K_{eq}$
The reaction will shift to the LEFT as it approaches equilibrium.
So $[\text{PCl}_5]$ will increase while $[\text{PCl}_3]$ and $[\text{Cl}_2]$ decrease.

Answer Shift LEFT

Nelson Chemistry 1chapter Review Answers

JE Gale



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