

Questions:

- 1) 5.0 moles of NH_3 and 2.0 moles of H_2 are introduced into a 4.00 L flask. At equilibrium, 2.5 moles of NH_3 gas remain. Calculate K_c for this reaction.



$$[\text{NH}_3] = \frac{5 \text{ mol} - 2.5 \text{ mol}}{4.00 \text{ L}} = 0.625 \text{ M}$$

$$[\text{H}_2] = \frac{2.0 \text{ mol}}{4.00 \text{ L}} = 0.5 \text{ M}$$

	H_2	N_2	NH_3
D	0.5	0	1.25
C	+3x	+x	-2x
E	0.5 + 3x	x	(1.25 - 2x)

- 2) The K_c value for the following equilibrium at 500°C is 49.



If 1.50 moles of H_2 and 1.50 moles of I_2 are introduced into a 2.00 L flask at 500°C , how many moles of HI are present at equilibrium?

$$[\text{H}_2] = \frac{1.5 \text{ mol}}{2 \text{ L}} = 0.75 \text{ M}$$

$$[\text{I}_2] = \frac{1.50 \text{ mol}}{2 \text{ L}} = 0.75 \text{ M}$$



$$0.75 - x$$

$$0.75 - x$$

$$+2x$$

$$0.75 - x$$

$$0.75 - x$$

$$2x$$

$$K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = \frac{(2x)^2}{(0.75-x)^2} = 49$$

- 3) What would happen to the values you calculated for K_c in the experiment if the Beer's Law constant (k) value you used was high (e.g. you used $6,000 \text{ M}^{-1}$ when the actual constant had a value of $5,000 \text{ M}^{-1}$)? Explain your answer.

Initial = 0, Run must be backward give concs

$$1.05 \text{ mol/L} \times 4 \text{ L} = 4.2 \text{ mol}$$

$$x = 0.58 \text{ M}$$

$$[\text{HI}] = 2x = 1.16 \text{ M}$$

$$\text{mol HI} =$$

$$\frac{1.16 \text{ mol/L} \times 2 \text{ L}}{1 \text{ L}} = 2.32 \text{ mol}$$

$$= 2.32 \text{ mol}$$

$$K_c = \frac{[\text{NH}_3]^2}{[\text{H}_2]^3 [\text{N}_2]}$$

$$= \frac{(0.625)^2}{(0.125)^3 (0.125)} = 0.421$$

General Chemistry Lab Manual 8th Edition Answers

Thomas Griffiths



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