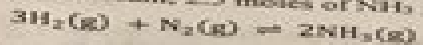


Questions:

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 the reaction.



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

$[\text{H}_2] = \frac{2.0 \text{ mol} - 0.75 \text{ mol}}{4.00 \text{ L}} = 0.3125 \text{ M}$

	H_2	N_2	NH_3
D	1 0.5	0	1.25
C	+3x = 0.9375	+x = 0.3125	-2x = 0.625
E	0.5 + 3x = 1.4375	x = 0.3125	(1.25 - 2x)

Initial = 0, Run must be backward give conv

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?

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

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

H₂ I₂ HI

H_2	+	I_2	\rightleftharpoons	$2\text{HI}(\text{g})$
0.75		0.75		0
-x		-x		+2x
0.75-x		0.75-x		2x

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

$7 = \frac{2x}{0.75-x}$
 $x = 0.58 \text{ M}$
 $[\text{HI}] = 2x = 1.16 \text{ M}$
 $1.16 \text{ mol} \times 2 \text{ L} = 2.32 \text{ mol}$

$K_c = \frac{[\text{NH}_3]^2}{[\text{H}_2][\text{N}_2]}$
 $(0.625)^2 = \frac{(1.4375)^2}{(0.3125)^2}$
 $= 0.421$

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.

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W.L.F. Armarego



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