(5) Initially 1.50 moles of N2(g) and 3.50 moles of H2(g) were added to a 1 L container at 700 °C. As a result of the reaction the equilibrium concentration of NH3(g) became 0.540 M.

(a) What is the value of the equilibrium constant for this reaction at the given temperature of 700 °C.

$X = \frac{0.540M}{2} = 0.270M$

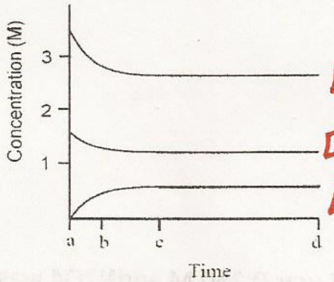
$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

I	1.50M	3.50M	0M
C	−X	−3X	+2X
E	1.23M	2.69M	0.540M

$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{[0.540M]^2}{[1.23M][2.69M]^3}$

$\boxed{K_c = 0.0122}$

In the above reaction we can monitor the change in concentration of reactants over time (just as we discussed when dealing with kinetics) and we can plot the data as follows:



$[H_2]$ (Large decrease)

$[N_2]$ (small decrease)

$[NH_3]$ (note the increase)

(b) Label each data plot as either[H2],[N2] or[NH3].

(c) Why does one data plot show an initial positive slope whereas the other two data plots show initial negative slopes?

The Forward reaction is initially faster than the reverse, resulting in a net decrease of the reactants of the Fwd Rxn & net increase of the Rev.

(d) Why does the uppermost plot have a steeper initial slope than the middle plot.

$[H_2]$ changes at a rate 3x faster than $[N_2]$ due to the 3:1 ratio of reactants

(e) At which time a, b, c or d is an equilibrium state reached? C

(ii) The equilibrium constant for this reaction was calculated as shown below. What is Kc for the reverse reaction?

$K_{rev} = \frac{[N_2][H_2]^3}{[NH_3]^2} = \frac{[1.23M][2.69M]^3}{[0.540M]^2}$ (Its the inverse of K_{fwd})