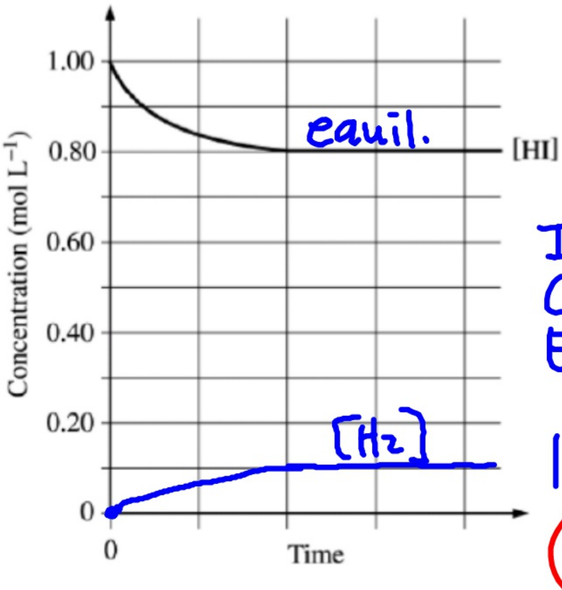
$$2\,HI(g) \rightleftarrows H_2(g) + I_2(g)$$

After a 1.0 mole sample of HI(g) is placed into an evacuated 1.0 L container at 700. K, the reaction represented above occurs. The concentration of HI(g) as a function of time is shown below.

(a) Write the expression for the equilibrium constant, Kc, for the reaction. $K_c = \frac{[H_2][I_2]}{[HI]^2}$ 1 pt



(b) What is [HI] at equilibrium?

$[HI] = 0.80\,M$ 1 pt

(c) Determine the equilibrium concentrations of $H_2(g)$ and $I_2(g)$.

	$2HI(g)$	$\rightleftarrows$	$H_2(g)$	+	$I_2(g)$
I	1 M		0 M		0 M
C	$-2x$		$+x$		$+x$
E	$1M - 2x$		x		x

$1M - 2x = 0.80\,M$ 1 pt

1 pt $x = 0.10\,M = [H_2] = [I_2]$

(d) On the graph above, make a sketch that shows how the concentration of $H_2(g)$ changes as a function of time.

(e) Calculate the value of the following equilibrium constants at 700. K.

(i) Kc

$$K_c = \frac{[H_2][I_2]}{[HI]^2} = \frac{(0.1)(0.1)}{(0.8)^2} = \boxed{0.016}$$ 1 pt

must agree w/ parts b & c

(ii) Kp

$$K_p = K_c\,(RT)^{\Delta n}$$

$$K_p = 0.016\,(0.0821 \times 700K)^0 = \boxed{0.016}$$ 1 pt

(f) At 1,000 K, the value of Kc for the reaction is 2.6×10^{-2}. In an experiment, 0.75 mole of HI(g), 0.10 mole of $H_2(g)$, and 0.50 mole of $I_2(g)$ are placed in a 1.0 L container and allowed to reach equilibrium at 1,000 K. Determine whether the equilibrium concentration of HI(g) will be greater than, equal, to, or less than the initial concentration of HI(g). Justify your answer.

1 pt

$$Q = \frac{[H_2][I_2]}{[HI]^2} = \frac{(0.10)(0.50)}{(0.75)^2} = 8.9 \times 10^{-2}$$

1 pt

$K < Q$ ∴ shift left ∴ [HI] increases