AP® Chemistry
2003 Sample Student Responses
Form B

1A, 1

CHEMISTRY

Section II

(Total time—90 minutes)

Part A

Time—40 minutes

YOU MAY USE YOUR CALCULATOR FOR PART A.

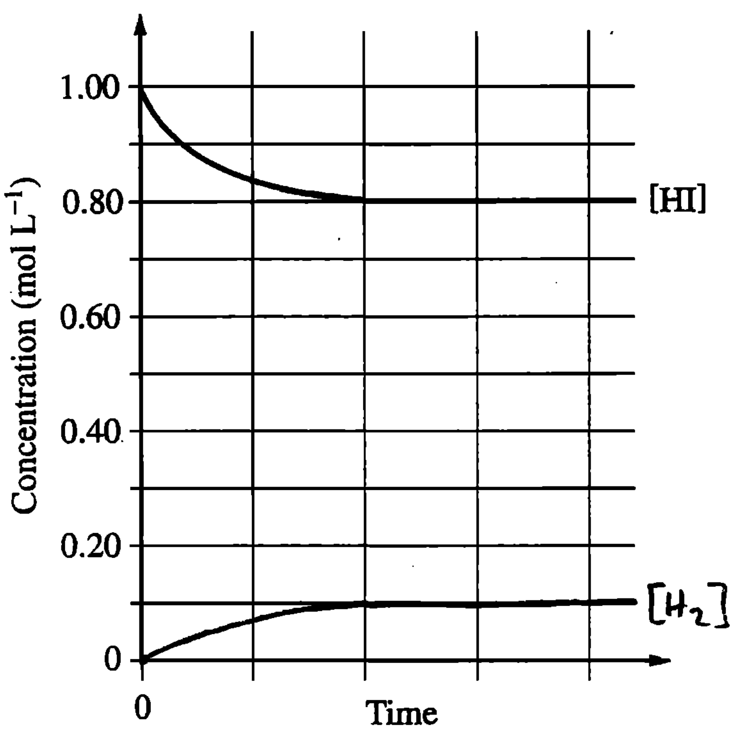
CLEARLY SHOW THE METHOD USED AND THE STEPS INVOLVED IN ARRIVING AT YOUR ANSWERS. It is to your advantage to do this, since you may obtain partial credit if you do and you will receive little or no credit if you do not. Attention should be paid to significant figures.

Be sure to write all your answers to the questions on the lined pages following each question in this booklet. Do NOT write your answers on the lavender insert.

Answer Question 1 below. The Section II score weighting for this question is 20 percent.

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

1. 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, K_c, for the reaction.

(b) What is [HI] at equilibrium?

1A 2

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

(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) K_c

(ii) K_p

(f) At 1,000 K, the value of K_c 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.

a) $K_c = \frac{[I_2][H_2]}{[HI]^2}$

b) $[HI] = 0.80\ M$

c) $n(H_2) = n(I_2) = n(HI) \div 2 = CV/2 = \frac{(0.20)(1.0)}{2} = 0.10\ mol$

$\Rightarrow [I_2] = [H_2] = \frac{n}{V} = \frac{0.10}{1.0} = \boxed{0.10\ M}$

e) i) $K_c = \frac{[0.10][0.10]}{[0.80]^2} = 0.015625 = \boxed{16. \times 10^{-3}}$

ii) $K_p = K_c (RT)^{\Delta n} = 0.015625 (0.0821 \times 700)^{(2-2)} = \boxed{16. \times 10^{-3}}$

f) $Q = \frac{(0.50)(0.10)}{(0.75)^2} = \frac{4}{45} = \boxed{8.9 \times 10^{-2}}$

$Q > K_c$ therefore the backward reaction is favoured to reach equilibrium $\Rightarrow$ [HI] will increase $\Rightarrow$ [HI] > initial

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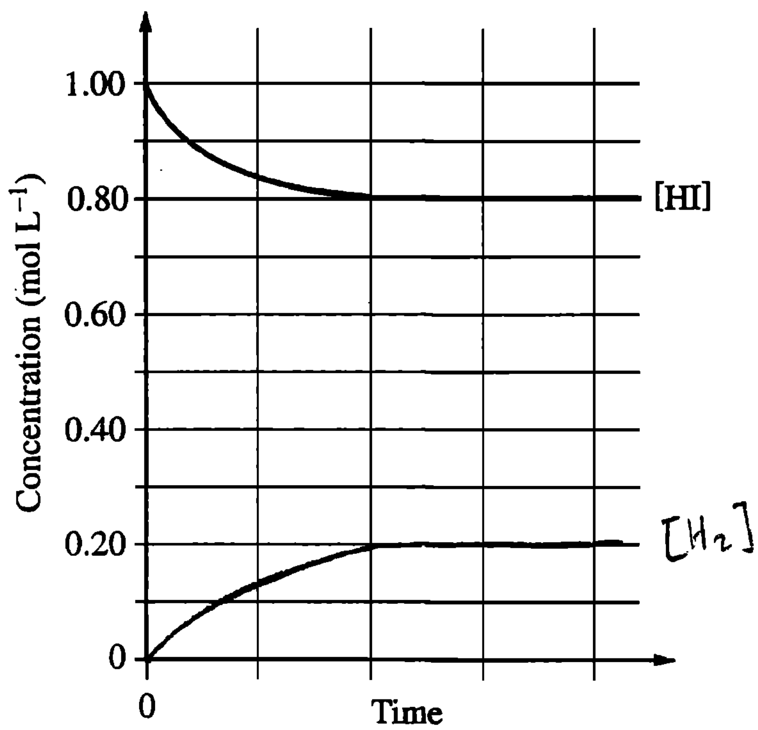

(a) Write the expression for the equilibrium constant, K_c, for the reaction.

(b) What is [HI] at equilibrium?

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1B2

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

(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) K_c

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(f) At 1,000 K, the value of K_c 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.

A) $K_c = \frac{[H_2][I_2]}{[HI]^2}$

B) Using the graph, the concentration of HI is 0.80 mol L^{-1}

c) $[H_2] = [I_2] = x$

$x = 1.0 \text{ mol } L^{-1} - 0.80 \text{ mol } L^{-1}$

$= 0.2 \text{ mol } L^{-1}$

e) $K_c = \frac{(0.2)^2}{(0.8)^2}$

$= 0.06$

$= 6 \times 10^{-2}$

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

$= 6 \times 10^{-2} (8.31 \cdot 700 K)^{2-2}$

$= 6 \times 10^{-2}$

f) $2HI_{(g)} \rightleftharpoons H_{2(g)} + I_{2(g)}$

$K_c = \frac{[H_2][I_2]}{[HI]^2}$

$= \frac{0.10 \times 0.50}{(0.75)^2}$

$= 0.089$

$= 8.9 \times 10^{-2}$

This is higher than the K_c given which means the ~~concentr~~ the reverse reaction is favored to increase the concentration of HI to keep the constant, 2.6×10^{-2}. Thus, the concentration of HI is greater than the initial concentration.

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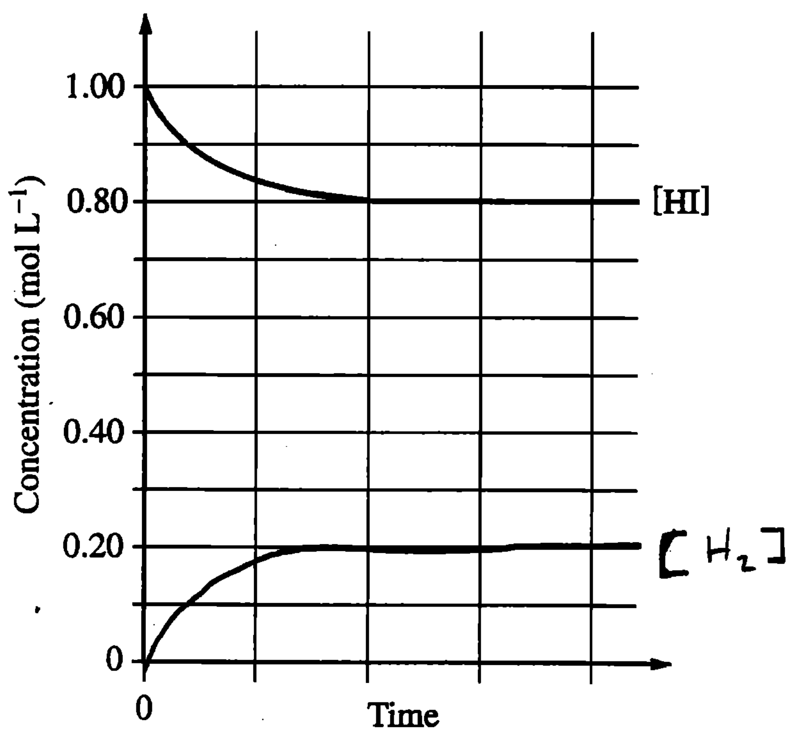

(a) Write the expression for the equilibrium constant, K_c, for the reaction.

(b) What is [HI] at equilibrium?

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1C2

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

(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) K_c

(ii) K_p

(f) At 1,000 K, the value of K_c 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.

a) $K_c = \frac{[H_2][I_2]}{[HI]}$

b) [HI] at equilibrium = 0.80 M

c) $2HI \rightleftharpoons H_2 + I_2$ x = 0.2 M

1.0M − x = 0.8M   x   x

therefore, $[H_2]$ and $[I_2]$ = 0.2 M at equilibrium.

e) (i) $K_c = \frac{[0.2M][0.2M]}{[1.0M - 0.2M]}$ = 0.05 = K_c

(ii) $K_p = K_c(RT)^{\Delta n}$ Δn = (2 mols) − (2 mols) = Ø

$K_p = (0.05)[(8.31\frac{J}{mol})(700K)]^0$ K_p = (0.05)(1) = 0.05 = K_p also

f) $2HI \rightleftharpoons H_2 + I_2$ $K_c = \frac{[H_2][I_2]}{[HI]}$ $2.6 \times 10^{-2} = \frac{x^2}{0.75 - x}$ (drop x, negligible)   x = 0.140 M

0.75 − x   x   x

This would decrease the molarity because 0.140 M would be subtracted from 0.75 M.

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