CollegeBoard
Advanced Placement Program

AP® Chemistry
2004 Sample Student Responses
Form B

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.

$$N_2(g) + 3\,H_2(g) \rightleftarrows 2\,NH_3(g)$$

1. For the reaction represented above, the value of the equilibrium constant, K_p, is 3.1×10^{-4} at 700. K.

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

(b) Assume that the initial partial pressures of the gases are as follows:

$p_{N_2} = 0.411$ atm, $p_{H_2} = 0.903$ atm, and $p_{NH_3} = 0.224$ atm.

(i) Calculate the value of the reaction quotient, Q, at these initial conditions.

(ii) Predict the direction in which the reaction will proceed at 700. K if the initial partial pressures are those given above. Justify your answer.

(c) Calculate the value of the equilibrium constant, K_c, given that the value of K_p for the reaction at 700. K is 3.1×10^{-4}.

(d) The value of K_p for the reaction represented below is 8.3×10^{-3} at 700. K.

$$NH_3(g) + H_2S(g) \rightleftarrows NH_4HS(g)$$

Calculate the value of K_p at 700. K for each of the reactions represented below.

(i) $NH_4HS(g) \rightleftarrows NH_3(g) + H_2S(g)$

(ii) $2\,H_2S(g) + N_2(g) + 3\,H_2(g) \rightleftarrows 2\,NH_4HS(g)$

1a) $K_p = \dfrac{P_{NH_3}^2}{P_{H_2}^3 \cdot P_{N_2}}$

b) i) $Q = \dfrac{P_{NH_3}^2}{P_{H_2}^3 \cdot P_{N_2}} = \dfrac{(0.224)^2}{(0.903)^3(0.411)} = 0.166$

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ii) At 700K, $K_p = 3.1 \times 10^{-4}$

$Q = 0.166$

Since $Q > K_p$ the reaction will proceed backwards.

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

$\Delta n = 2 - 4 = -2$

$K_p = K_c(RT)^{-2}$

$K_c = K_p(RT)^2 = (3.1 \times 10^{-4})[(0.0821)(700)]^2$

$K_c = \underline{\underline{1.0}}$

d) i) ~~$K_p = 1.8 \times 3$~~ $K_p = \dfrac{1}{8.3 \times 10^{-3}} = \underline{\underline{1.2 \times 10^2}}$

ii) The given reaction is the sum of the following two reactions

① $2NH_{3(g)} + 2H_2S \rightleftharpoons 2NH_4HS$

② $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_3$

K_p for reaction ① $= (8.3 \times 10^{-3})^2 = 6.889 \times 10^{-5}$

K_p for reaction ② $= 3.1 \times 10^{-4}$

K_p for the overall reaction $= K_p$ for reaction ① $\times$ K_p for reaction ②

$= 6.889 \times 10^{-5} \times 3.1 \times 10^{-4}$

$= \underline{\underline{2.1 \times 10^{-8}}}$

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1 B 1

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A) $K_p = \dfrac{P_{NH_3}^{2}}{P_{N_2}\,P_{H_2}^{3}}$

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ADDITIONAL PAGE FOR ANSWERING QUESTION 1.

B) i) Q is the initial constant before equilibrium has been reached

Thus, $Q = \frac{P_{NH_3}^2}{P_{N_2} P_{H_2}^3} = \frac{(.224)^2}{(.411)(.903)^3} = \boxed{.166}$ $\boxed{Q = .166}$

no units for Q

ii) Q is .166, whereas, Kp is 3.1×10^{-4}. Thus, Q is bigger than Kp. Therefor, in order to reach equilibrium (Kp's value), the reaction must go left. Going left (producing more reactants) will decrease the value of Q until it reaches 3.1×10^{-4} (Kp's value)

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

R is a constant

T is the temp. in kelvin.

N is the diff. in number of gaseus products to gaseus reactants.

$3.1 \times 10^{-4} = K_c(8.31 \times 700)^{(2-4)}$

$K_c = 10489.62 \approx 10,000 = \boxed{1.0 \times 10^4} = K_c$

D) i) The ~~opposite~~ reverse reaction for equilibrium, results in the original constant changed to its reciprical.

Thus, $8.3 \times 10^{-3} \rightarrow \frac{1}{8.3 \times 10^{-3}} = 120.5 \approx \boxed{1.2 \times 10^2}$

ii) This reaction is the combination of $N_2 + 3H_2 \rightleftharpoons 2NH_3$ & ($NH_3 + H_2S \rightleftharpoons NH_4HS$ multiplied by 2). Thus we can obtain the answer by multiplying the constants because we know the constants for each ~~step~~ reaction

$(3.1 \times 10^{-4})\{2(8.3 \times 10^{-3})\}^2 = \boxed{5.1 \times 10^{-6}}$

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   (ii) $2\,H_2S(g) + N_2(g) + 3\,H_2(g) \rightleftarrows 2\,NH_4HS(g)$

(a) $K_p = \frac{[NH_3]^2}{[N_2][H_2]^3}$

(b) (i) $Q = \frac{(0.224)^2}{0.411 \times (0.903)^3} = 0.166$

(ii) Since Q, the reaction quotient's value is less than 1, this reaction will proceed from left to right spontaneously at the temperature of 700K.

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ADDITIONAL PAGE FOR ANSWERING QUESTION 1. 1C2

(c) $K_p = K_c (RT)^{\Delta n}$, $3.1 \times 10^{-4} = K_c (0.0821 \times 700)^{-2}$

$\therefore K_c = \frac{3.1 \times 10^{-4}}{(0.0821 \times 700)^{-2}} = 10.2$

(d) ① $K_p = \frac{[NH_3][H_2S]}{[NH_4HS]}$ = Reciprocal value of the original reaction.

$\therefore 1/8.3 \times 10^{-3} = 1.2 \times 10^2$

② $K_p = \frac{[NH_4HS]^2}{[H_2S]^2[N_2][H_2]^3}$ = (original reaction's K_p)^2 × (the very first reaction <$N_2 + 3H_2 \leftrightarrow 2NH_3$>'s K_p)

$\therefore (8.3 \times 10^{-3})^2 \times 3.1 \times 10^{-4} = 2.1 \times 10^{-8}$

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