Key 3A white

5. The reaction below was studied at 250°C.

$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$

At equilibrium, the partial pressures of the gases are as follows: $PCl_5 = 2.0 \times 10^{-2}$ atm, $PCl_3 = 4.2 \times 10^{-2}$ atm, and $Cl_2 = 4.2 \times 10^{-4}$ atm. What is the value of K_p for the reaction?

(3)

$$K_p = \frac{P_{PCl_3} P_{Cl_2}}{P_{Cl_5}} = \frac{(4.2 \times 10^{-2})(4.2 \times 10^{-4})}{2.0 \times 10^{-2}} = 8.8 \times 10^{-4}$$

6. At 2010 K, the equilibrium constant, K_c, for the following reaction is 4.0×10^{-4}.

$N_2(g) + O_2(g) \rightleftharpoons 2\,NO(g)$

If the equilibrium concentrations of N_2 and O_2 are 0.25 mol/L and 0.33 mol/L, what is the equilibrium concentration of NO?

(4)

$N_2 + O_2 \rightleftharpoons 2NO$ $\quad K_c = \frac{[NO]^2}{[N_2][O_2]}$

I

C

E 0.25 0.33 x $\quad 4.0 \times 10^{-4} = \frac{x^2}{[0.25][0.33]}$

$x = 5.7 \times 10^{-3}$

7. We place 0.064 mol $N_2O_4(g)$ in a 4.00 L flask at 200. After reaching equilibrium, the concentration of $NO_2(g)$ is 0.0030 M. What is K_c for the reaction below?

(4)

$N_2O_4(g) \rightleftharpoons 2\,NO_2(g)$

I 0.016 M 0

C $-x$ $+2x$

E $0.016 - x$ $2x$

$2x = 0.0030$ M

$x = 0.0015$ M

$0.016 - 0.0015 = 0.0145$

$K_c = \frac{[0.0030]^2}{0.0145}$

$K_c = 6.2 \times 10^{-4}$

8. A mixture of 0.200 mol NO and 0.200 mol CO_2 is placed in a 1.00 L flask and allowed to reach equilibrium at a given temperature. Analysis of the equilibrium mixture indicates that 0.067 mol of CO is present. Calculate K_c for the reaction.

$NO(g) + CO_2(g) \rightleftharpoons NO_2(g) + CO(g)$

(4)

I 0.200 0.200 0 0

C $-x$ $-x$ $+x$ $+x$

E $0.200 - x$ $0.200 - x$ x x

$x = 0.067$

$0.200 - 0.067 = 0.133$

$K_c = \frac{[0.067]^2}{[0.133]^2}$

$K_c = 0.25$

CHM152 Quiz 3a Spring 2019 25 Pts Name: Key
Show all work to receive credit.

1. Write the expression for K_c for the reaction below.

$Mg_3(PO_4)_2(s) \rightleftharpoons 3\ Mg^{2+}(aq) + 2\ PO_4^{3-}(aq)$

(2) $K_c = [Mg^{2+}]^3 [PO_4^{3-}]^2$

2. Write a balanced chemical reaction which corresponds to the following equilibrium constant expression.

$K_p = \dfrac{P_{NOBr}}{P_{NO} P_{BR_2}^{1/2}}$

(2) $NO(g) + \frac{1}{2} Br_2(g) \rightleftharpoons NOBr(g)$

3. For the following reaction,

1 $2\ SO_3(g) \rightleftharpoons 2\ SO_2(g) + O_2(g)$ $K_{P_1} = \dfrac{P_{SO_2}^2 P_{O_2}}{P_{SO_3}^2}$

, the equilibrium constant, K_p, is 1.32 at 627°C. What is the equilibrium constant for the reaction below?

(3) 2 $SO_3(g) \rightleftharpoons SO_2(g) + 1/2\ O_2(g)$ $K_{P_2} = \dfrac{P_{SO_2} P_{O_2}^{1/2}}{P_{SO_3}}$

note $K_{P_2} = \sqrt{K_{P_1}} = 1.15$

4. The oxidation of sulfur dioxide produces sulfur trioxide.

$2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g)$

Calculate the value of K_p, given that K_c for the reaction is 2.3×10^4 at 999 K. ($R = 0.08206$ L·atm/mol·K)

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

$= 2.3 \times 10^4 (0.0821 \cdot 999)^{-1} = 280$ or 2.8×10^2

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