

Unit 10: Equilibrium Practice 1**Calculating K_c and K_p**

1. Consider the following chemical reaction:



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

$$K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$$

- (b) At equilibrium in a particular experiment, the concentrations of
- H_2
- ,
- I_2
- , and
- HI
- were 0.15 M, 0.033 M, and 0.55 M, respectively. Calculate the value of the equilibrium constant,
- K_c
- , for the reaction.

$$K_c = \frac{(0.55)^2}{(0.15)(0.033)} = 61$$

2. Acetic acid is a weak acid that dissociates into the acetate ion and a proton in aqueous solution:



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

$$K_c = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]}$$

- (b) At equilibrium at 25 °C, a 0.100 M solution of acetic acid has the following concentrations:
- $[\text{CH}_3\text{COOH}] = 0.0990 \text{ M}$
- ,
- $[\text{CH}_3\text{COO}^-] = 1.33 \times 10^{-3} \text{ M}$
- , and
- $[\text{H}^+] = 1.33 \times 10^{-3} \text{ M}$
- . Calculate the value of the equilibrium constant,
- K_c
- , for the reaction.

$$K_c = \frac{[1.33 \times 10^{-3}][1.33 \times 10^{-3}]}{[0.0990]} = 1.79 \times 10^{-5}$$

3. At 22 °C,
- $K_p = 0.070$
- for the equilibrium:



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

$$K_p = P_{\text{NH}_3} \times P_{\text{H}_2\text{S}}$$

- (b) A sample of solid
- NH_4HS
- is placed in a closed vessel and allowed to equilibrate. Calculate the equilibrium partial pressure (atm) of ammonia, assuming that some solid
- NH_4HS
- remains.

$$K_p = P_{\text{NH}_3} \times P_{\text{H}_2\text{S}} = 0.070 = x^2$$

$$0.26 \text{ atm} = x$$

4. Given the following reaction at equilibrium at 450.0 °C:



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

$$K_p = P_{\text{CO}_2}$$

- (b) If
- $P_{\text{CO}_2} = 0.0160 \text{ atm}$
- , calculate the value of
- K_c
- .

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

$$0.0160 = K_c [(0.0821)(450.0 + 273)]^{(1 - 0)}$$

$$\frac{0.0160}{(59.36)} = K_c \frac{(59.36)}{(59.36)}$$

$$2.70 \times 10^{-4} = K_c$$

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Ian Pickup



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