Calculating the Equilibrium Constant For A Reaction

Name ______________________ Section ______________

1'. Write the equilibrium expression, given the following general equation.

$$aA_{(g)} + bB_{(g)} \rightleftharpoons cC_{(g)} + dD_{(g)}$$

$$\mathbf{K = \frac{[C]^c[D]^d}{[A]^a[B]^b}}$$

Distinguish between the K_p and K_c for the reaction.

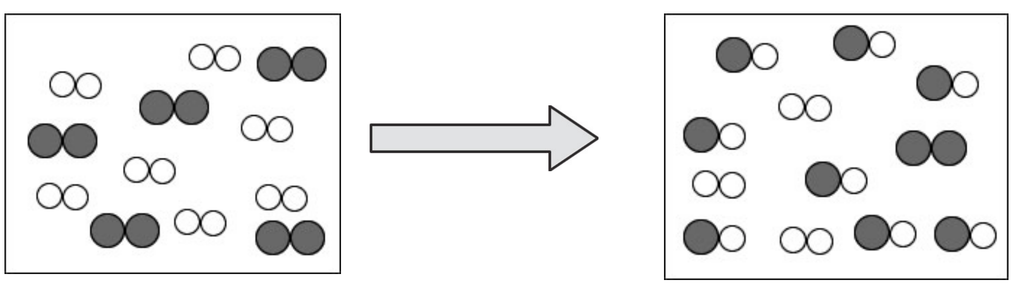
$$\mathbf{K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}}$$

$$\mathbf{K_p = \frac{P_C^c P_D^d}{P_A^a P_B^b}}$$

K_c is the equilibrium constant in which the concentration of reactant and products are expressed in terms of molarity. K_p is the equilibrium constant in which the amounts of reactant and products are expressed in terms of partial pressures measured in atm or mm Hg.

1. Below are listed three reactions. Associated with each reaction is a 1.0 L container with a particulate level representation of the reaction before the reaction has occurred. To the right is the 1.0 L container with a particulate level representation of the reaction after attaining equilibrium. In each case indicate whether you think the equilibrium constant for the reaction is greater than 1, less than 1, or equal to 1. In each case support your answer with a brief explanation.

   a. Reaction I: $A_{2(g)} + B_{2(g)} \rightleftharpoons 2AB_{(g)}$ (where $A_{2(g)}$  and $B_{2(g)}$ is ∞)

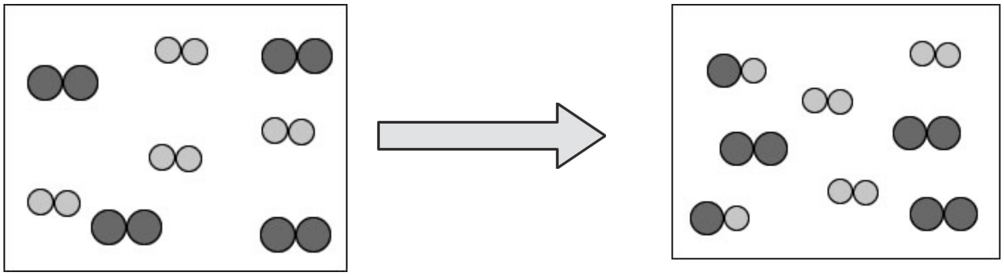


Container before the reaction started

Container after the reaction achieves equilibrium

$$\mathbf{K = \frac{[AB]^2}{[A_2]^1[B_2]^1} = \frac{[8]^2}{[3]^1[1]^1} = 7.1}$$

b. Reaction II: $C_{2}(g) + D_{2}(g) \rightleftharpoons 2CD(g)$ (where $C_{2}(g)$  and $D_{2}(g)$ is )

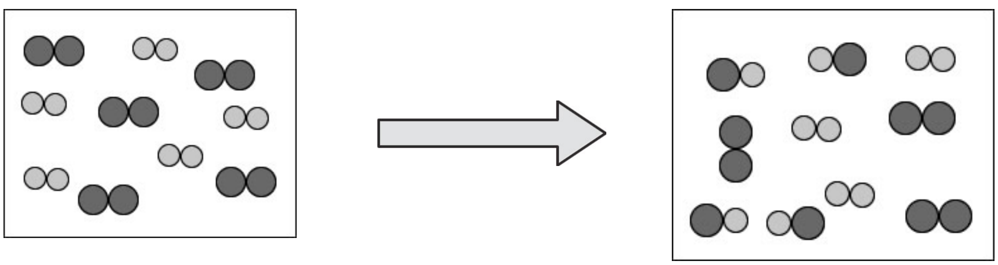


Container before the reaction started

Container after the reaction achieves equilibrium

$$\mathbf{K = \frac{[CD]^{2}}{[C_{2}]^{1}[D_{2}]^{1}} = \frac{[2]^{2}}{[3]^{1}[3]^{1}} = 0.44}$$

c. Reaction II: $X_{2}(g) + Y_{2}(g) \rightleftharpoons 2XY(g)$ (where $X_{2}(g)$  and $Y_{2}(g)$ is )

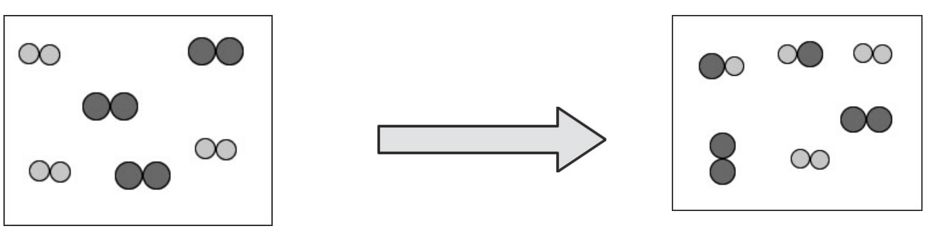


Container before the reaction started

Container after the reaction achieves equilibrium

$$\mathbf{K = \frac{[XY]^{2}}{[X_{2}]^{1}[Y_{2}]^{1}} = \frac{[4]^{2}}{[3]^{1}[3]^{1}} = 1.8}$$

d. If any of the cases (K > 1, K < 1, or K = 1) did not appear in the three examples above, use the space below to draw the before container, and the equilibrium container for the missing case.



Container before the reaction started

Container after the reaction achieves equilibrium

Explain how your model properly represents the particular case.

Assume the reaction is

$$G_2(g) + J_2(g) \rightleftharpoons 2JG(g)$$

$$\mathbf{K = \frac{[JG]^2}{[G_2]^1[J_2]^1} = \frac{[2]^2}{[2]^1[2]^1} = 1}$$

NOTE: It would be an unusual for a reaction to have an equilibrium constant equal to 1.

2'. Distinguish between the equilibrium constant and the equilibrium constant expression for a chemical reaction.

The equilibrium constant is a number whose value reflects whether the amount of products is greater or less than the amount of reactants at equilibrium. When K is >1, the amount of products is greater than reactants; the reaction favors products. When K is <1, the amount of products is less than reactants; the reaction favors reactants.

The equilibrium constant expression is the quotient of product equilibrium concentration to reactant equilibrium concentrations as determined from the balanced chemical equation.

2. The following reaction is at equilibrium at a particular temperature

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

and the $[H_2]_{eq} = 0.012$ M, $[I_2]_{eq} = 0.15$ M and $[HI]_{eq} = 0.30$ M. Calculate the magnitude of K_c for the reaction.

The equilibrium constant expression for the reaction is

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

Substituting the equilibrium concentrations into the equilibrium expression,

$$\mathbf{K_c = \frac{(0.30)^2}{(0.012)(0.15)} = 50}$$

The magnitude of the equilibrium constant is 50 for the temperature at which the data was collected.

3. Using the equilibrium constant calculated in 2, calculate the magnitude of the equilibrium constant for the following reactions at the same temperature.

i) $2HI(g) \rightleftharpoons H_2(g) + I_2(g)$

The equilibrium constant expression from part b is

$$K_c = \frac{[HI]^2}{[H_2][I_2]} = 50$$

When the reaction is reversed the new equilibrium constant expression becomes:

$$K_c' = \frac{[H_2][I_2]}{[HI]^2} = \frac{1}{K_c} = \frac{1}{50} = 0.02$$

ii) $\frac{1}{2} H_{2(g)} + \frac{1}{2} I_{2(g)}$ ä $HI_{(g)}$

The equilibrium constant expression from part b is

$$K_c = \frac{[HI]^2}{[H_2][I_2]} = 50$$

When the coefficients in the reaction are changed, the new equilibrium constant expression becomes;

$$K_c'' = \frac{[HI]}{[H_2]^{1/2}[I_2]^{1/2}} = \sqrt{K_c} = 7.07$$

4. The initial concentration of both H_2 and I_2 is 0.250 M. The reaction occurs as shown below,

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

When equilibrium is achieved the concentration of HI is 0.393 M. Calculate the magnitude of K_c for the reaction.

	$H_{2(g)}$	+	$I_{2(g)}$	$\rightleftharpoons$	$2HI_{(g)}$
Initial	**0.250 M**		**0.250 M**		**0**

We know the reaction proceeds from left to right because there is no HI present initially and the equilibrium amount of HI is 0.393 M. To form HI some H_2 and some I_2 must react. How much? To determine how much we must use the coefficients from the balanced chemical equation.

The amount of H_2 reacting is,

$0.393\text{ M HI}\left(\frac{1\text{ mol }H_2}{2\text{ mol HI}}\right) = 0.196\text{ M }H_2$ Note: that this is also the amount of I_2.

The equilibrium amount of H_2 and I_2 is determined by subtracting the amount reacting from the initial amount.

$[H_2]_{eq} = 0.250\text{ M} - 0.196\text{ M} = 0.054\text{ M}$
$[I_2]_{eq} = 0.250\text{ M} - 0.196\text{ M} = 0.054\text{ M}$

$$K_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{[0.393\text{ M}]^2}{[0.054\text{ M}][0.054\text{ M}]} = 53.0$$

5. A vessel initially has a partial pressure of NO equal to 0.526 atm and a partial pressure of Br_2 equal to 0.329 atm. At equilibrium the partial pressure of Br_2 is 0.203 atm. Calculate K_p for the reaction ($K_p = 0.475$)

$$2NO_{(g)} + Br_{2(g)} \rightleftharpoons 2NOBr_{(g)}$$

	$2NO_{(g)}$	+	$Br_{2(g)}$	$\rightleftharpoons$	$2NOBr_{(g)}$	
Initial	**0.526 atm**		**0.329 atm**		**0**	**let $x = P_{Br_2(reacting)}$**
Change	**-2x**		**-x**		**+2x**	
Equilibrium	**0.526 - 2x**		**0.329 - x**		**0 + 2x**	

$[Br_2]_{eq} = 0.203\text{ atm} = 0.329\text{ atm} - x$
$x = 0.329\text{ atm} - 0.203\text{ atm} = 0.126\text{ atm}$
$[NO]_{eq} = 0.526 - 2x = 0.526 - 2(0.126\text{ atm}) = 0.274\text{ atm}$
$[NOBr]_{eq} = 2x = 2(0.126\text{ atm}) = 0.252\text{ atm}$

$$K_p = \frac{P^2_{NOBr}}{P^2_{NO} \cdot P_{Br_2}} = \frac{(0.252)^2}{(0.274)^2 \cdot (0.203)} = 4.17$$