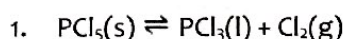


Totally Epic AP Chem Review: Equilibrium in a Day!

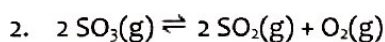
EQUILIBRIUM CONSTANT EXPRESSIONS: The product concentrations appear in the numerator and the reactant concentrations in the denominator. Each concentration is raised to the power of its stoichiometric coefficient in the balanced equation.

- K_c is for concentration (aqueous)
- K_p is for partial pressure (gases)
- "K" values are written without units
- Pure solids and pure liquids do not appear in expression

Practice: Write the expressions for K_c and K_p for the following process:



$$K_c = [\text{Cl}_2], \quad K_p = P_{\text{Cl}_2}$$



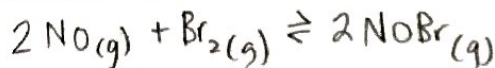
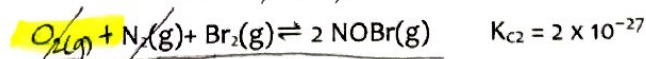
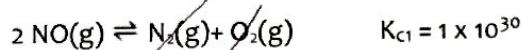
$$K_c = \frac{[\text{SO}_2]^2 [\text{O}_2]}{[\text{SO}_3]^2}, \quad K_p = \frac{(P_{\text{SO}_2})^2 (P_{\text{O}_2})}{(P_{\text{SO}_3})^2}$$

Manipulating Reactions

- **Stoichiometric Coefficients:** If you multiply the coefficients in the equation by a factor, the K is raised to the power of the multiplication factor. 2x is K^2 , 3x is K^3 , etc
- **Reversing Equations:** take the reciprocal of K ($1/K$)
- **Adding Equations:** multiply respective Ks ($K_1 \times K_2 \times K_3 \dots$)

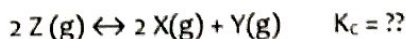
Practice:

1. Calculate the value of K_c for the reaction $2 \text{NO}(\text{g}) + \text{Br}_2(\text{g}) \rightleftharpoons 2 \text{NOBr}(\text{g})$ using the following information.



$$K_c = K_{c1} \times K_{c2} = (1 \times 10^{30})(2 \times 10^{-27}) = \boxed{2000}$$

2. For the reaction $2 \text{X}(\text{g}) + \text{Y}(\text{g}) \leftrightarrow 2 \text{Z}(\text{g})$, $K_c = 4.0 \times 10^4$. Determine the value of the equilibrium constant, K_c , for the following reaction:



a. 2.5×10^{-5}

b. 2.5×10^{-4}

c. 4.0×10^{-5}

d. 4.0×10^{-4}

Reverse! $K_{\text{new}} = \frac{1}{K} = \frac{1}{4 \times 10^4} = \frac{1}{4} \times 10^{-4}$
 $= 0.25 \times 10^{-4} = 2.5 \times 10^{-5}$

THE REACTION QUOTIENT, Q : When you need to know the answer to the question, "Is the system at equilibrium?"

A: The answer can be yes or no !

For the general reaction: $aA + bB \rightleftharpoons cC + dD$

$$Q = \frac{[C]^c [D]^d}{[A]^a [B]^b} \quad \text{or} \quad \frac{(P_C)^c (P_D)^d}{(P_A)^a (P_B)^b}$$

→ Reminder: Q has the appearance of K but the concentrations are not necessarily at equilibrium!

What does Q mean?

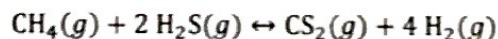
1. If $K > Q$, system not at equilibrium: forward reaction is favored (shift right) to make $Q = K$.
2. If $K = Q$, the system is at equilibrium.
3. If $K < Q$, system not at equilibrium: reverse reaction is favored (shift left) to make $Q = K$.

Practice:

1. The value of the equilibrium constant, K_c , at 25°C is 8.1 for the following reaction: $2 \text{SO}_3(\text{g}) \rightleftharpoons 2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g})$. What must happen for the reaction to reach equilibrium if the initial concentrations of all three species was 2.0 M?
 - a. The rate of the forward reaction would increase, and $[\text{SO}_3]$ would decrease.
 - b. The rate of the reverse reaction would increase, and $[\text{SO}_3]$ would decrease.
 - c. Both the rate of the forward and reverse reactions would increase, and the value for the equilibrium constant would also increase.
 - d. No change would occur in either the rate of reaction or the concentrations of any of the species.

$$Q = \frac{[\text{SO}_2]^2 [\text{O}_2]}{[\text{SO}_3]^2} = \frac{(2)^2 (2)}{(2)^2} = 2 \quad \left. \begin{array}{l} K > Q \\ 8.1 > 2 \end{array} \right\}$$

2. Consider the following reaction:



1.00 M CH_4 , 1.00 M CS_2 , 2.00 M H_2S and 2.00 M H_2 are mixed in a reaction vessel at 960°C . At this temperature, the reaction will make more methane and hydrogen sulfide gases. What is a possible K for this reaction?

- a. $K = 16$
- b. $K = 8$

- c. $K = 4$
- (d) $K = 1$

⇒ shifts left

$K < Q$

$K < 4$

$$Q = \frac{[\text{CS}_2] [\text{H}_2]^4}{[\text{CH}_4] [\text{H}_2\text{S}]^2} = \frac{(1)(2)^4}{(1)(2)^2} = 2^2 = 4$$

Calculating with the Equilibrium Expression:

You MUST have a balanced equation.

If the amounts are given in moles BE WARY – you must convert to M (molarity).

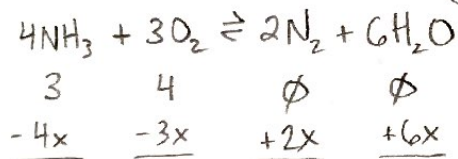
- Write the Equilibrium Constant Expression for K_c or K_p
- Set up RICE TABLE
 - R = Balanced Reaction
 - I = Initial concentrations
 - C = Change in concentration Remember: Everything changes stoichiometrically!
 - E = Equilibrium concentrations These are the concentrations (pressures) of all species at equilibrium

Hint: If none of the initial concentrations are zero, then Q must be calculated first to determine the direction of the shift (who gains and loses) before calculating the equilibrium concentrations.

Practice:

- Ammonia and oxygen react according to the following equilibrium: $4 \text{NH}_3(\text{g}) + 3 \text{O}_2(\text{g}) \rightleftharpoons 2 \text{N}_2(\text{g}) + 6 \text{H}_2\text{O}(\text{g})$.
A 1.0 liter flask is initially filled with 4.0 mol of oxygen and 3.0 mol of ammonia, and 1.0 mol of N_2 is present at equilibrium. How much oxygen is present at equilibrium?

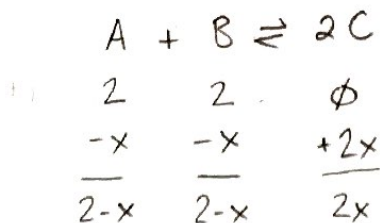
- a. 1.0 mol O_2 b. 1.5 mol O_2 **c.** 2.5 mol O_2 d. 3.0 mol O_2



$$\begin{array}{c}
 1.0 \\
 \Rightarrow x = 0.5 \Rightarrow [\text{O}_2] = 4 - 3x = 4 - 3(0.5) = 4 - 1.5 = 2.5
 \end{array}$$

- Here is a general reaction with a K value of 16: $\text{A}(\text{aq}) + \text{B}(\text{aq}) \rightleftharpoons 2 \text{C}(\text{aq})$.
Initially, $[\text{A}] = [\text{B}] = 2.0 \text{ M}$. Solve for the equilibrium concentration of each substance.

- a. $[\text{A}] = [\text{B}] = 0.67 \text{ M}$, $[\text{C}] = 1.3 \text{ M}$ **c.** $[\text{A}] = [\text{B}] = 0.67 \text{ M}$, $[\text{C}] = 2.7 \text{ M}$
b. $[\text{A}] = [\text{B}] = 1.6 \text{ M}$, $[\text{C}] = 0.88 \text{ M}$ d. $[\text{A}] = [\text{B}] = 0.50 \text{ M}$, $[\text{C}] = 3.0 \text{ M}$

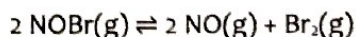


$$K_c = \frac{[\text{C}]^2}{[\text{A}][\text{B}]} = \frac{(2x)^2}{(2-x)^2} = 16$$

$$\begin{array}{l}
 [\text{A}] = [\text{B}] = 2 - 1.33 = 0.67 \\
 [\text{C}] = 2(1.33) = 2.66
 \end{array}$$

$$\begin{array}{l}
 \Rightarrow \frac{2x}{2-x} = 4 \quad 2x = 4(2-x) = 8 - 4x \\
 6x = 8 \Rightarrow x = \frac{8}{6} = 1.33
 \end{array}$$

- The reaction below came to equilibrium at a temperature of 100°C . At equilibrium the partial pressure due to NOBr was 4 atm, the partial pressure due to NO was 4 atm, and the partial pressure due to Br_2 was 2 atm. What is the equilibrium constant, K_p , for this reaction at 100°C ?



- a. $\frac{1}{4}$ b. $\frac{1}{2}$ c. 1 **d.** 2

$$K_p = \frac{(P_{\text{NO}})^2 (P_{\text{Br}_2})}{(P_{\text{NOBr}})^2} = \frac{(4)^2 (2)}{(4)^2} = 2$$