Al-Mustaqbal-College University

Chemical Engineering and Petroleum Industry Department

Analytical chemistry

First class / first term

Lecture Three part 2

By

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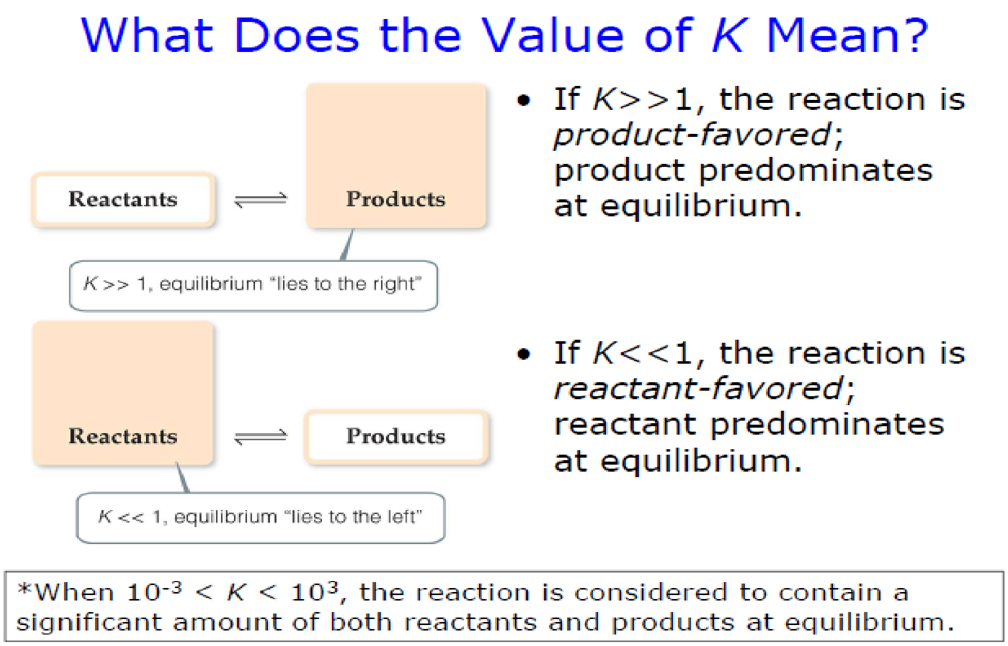
Lecture three

3.4 The Equilibrium Constant

For a reaction: aA + bB $\leftrightarrow$ cC + d

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

The **equilibrium constant, Kc,** is the ratio of the equilibrium concentrations of products over the equilibrium concentrations of reactants each raised to the power of their stoichiometric coefficients.



Note : if K > 1 the reaction is forward (product to right)
If K < 1 the reaction is reverse (reactant to left)

3.4.1Calculating Equilibrium Concentrations

Use ICE Tables To Solve Equilibrium Problems For Kc Or Equilibrium Amounts:

1. I = initial concentration: Initial concentration of reactants are usually given; initial [Product]'s are assumed to be 0 unless otherwise specified.

2. C = change in concentration: Assign change as the variable x; use the stoichiometry of the reaction to assign changes for all species.

3. E = equilibrium concentration: E = I + C

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Example 1\\ In an analysis of the following reaction at 100°C

$$Br_{2(g)} + Cl_{2(g)} \leftrightarrow 2BrCl_{(g)}$$

the equilibrium concentrations are $[Br_2]= 2.3\times10^{-3}$M,$[Cl_2] = 1.2 \times 10^{-2}$M, $[BrCl] = 1.4 \times 10^{-2}$ M. Write the equilibrium expression and calculate Kc for this reaction.

Solution:

$$K_c = \frac{[BrCL]^2}{[Br]_2[Cl]_2}$$

$$K_c = \frac{[1.4\times10^{-2}]^2}{[2.3\times10^{-3}][1.2\times10^{-2}]} = 7.1$$

Example 2\\ **The chemicals A and B react as follows to produce C and D:**

A + B $\leftrightarrow$ C + d , $K = \frac{[D][C]}{[A][B]}$, The equilibrium constant K has a value of 0.30. Assume 0.20 mol of A and 0.50 mol of B are dissolved in 1.00 L, and the reaction proceeds. Calculate the concentrations of reactants and products at equilibrium.

Solve :

	[A]	[B]	[C]	[D]
Initial	0.20	0.50	0	0
Change (x = mmol/mL reacting)	$-x$	$-x$	$+x$	$+x$
Equilibrium	$0.20 - x$	$0.50 - x$	x	x

We can substitute these values in the equilibrium constant expression and solve for x:

$$K = \frac{[D][C]}{[A][B]}$$

$$0.3 = \frac{(x)(x)}{(0.2-x)(0.5-x)}$$

$$x^2 = (0.10 - 0.70x + x^2)0.30$$

$$0.70x^2 + 0.21x - 0.030 = 0$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$= \frac{-0.21 \pm \sqrt{(0.21)^2 - 4(0.70)(-0.030)}}{2(0.70)}$$

$$= \frac{-0.21 \pm \sqrt{0.044 + 0.084}}{1.40} = 0.11\,M$$

[A] = 0.20 − x = 0.09M

[B] = 0.50 − x = 0.39M

[C] = [D] = x = 0.11M

Example 3 Determine the initial concentration of HI if the initial concentrations of H_2 and I_2 are both 0.10 *M* and their equilibrium concentrations are both 0.043 *M* at 430°C. The value of Kc = 54.3

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

Solution:

	$H_2(g)$ +	$I_2(g)$ $\rightleftharpoons$	$2HI(g)$
Initial	0.10	0.10	y
Change	-x	-x	+2x
Equilibrium	0.043	0.043	y + 2x

First solve for x: 0.10 – x = 0.043; x = 0.057

Then solve for y: at equilibrium we have

$$K_c = 54.3 = \frac{[HI]^2}{[H_2][I_2]} = \frac{(y + 0.114)^2}{(0.043)(0.043)}$$

$$(y + 0.114)^2 = (54.3)(0.043)^2 \;\Rightarrow\; y = \sqrt{(54.3)(0.043)^2} - 0.114 = 0.203$$

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Example 4 A closed system initially containing 1×10^{-3} M H_2 and 2×10^{-3} M I_2 at 448℃ is allowed to reach equilibrium. Analysis of the equilibrium mixture shows that the concentration of HI is 1.87×10^{-3} M. Calculate Kc at 448℃ for the reaction taking place, which is : $H_{2(g)} + I_{2(g)} \leftrightarrow 2HI_{(g)}$

solve :

1)

	H_2	I_2	2HI
initial	$1 * 10^{-3}$	$2 * 10^{-3}$	0
change			
equilibrium			$1.87 * 10^{-3}$

2)

	H_2	I_2	2HI
initial	$1*10^{-3}$	$2*10^{-3}$	0
change	-X	-X	+2X
equilibrium	$1*10^{-3}$ –X	$1*10^{-3}$ –X	$1.87*10^{-3}$

$0+2X = 1.87 \times 10^{-3} \rightarrow X = 9.35 * 10^{-4}$ M = [HI]

3)

	H_2	I_2	2HI
initial	$1 * 10^{-3}$	$2 * 10^{-3}$	0
change	$-9.35 * 10^{-4}$	$-9.35 * 10^{-4}$	$1.87 * 10^{-3}$
equilibrium	$1*10^{-3} - 9.35 * 10^{-4}$	$1.87 * 10^{-3} - 9.35 * 10^{-4}$	$1.87 * 10^{-3}$

$[H_2] = 1*10\text{-}3 - 9.35 * 10^{-4} = 6.5 * 10^{-5}$ M

$[I_2] = 1.87 * 10^{-3} - 9.35 * 10^{-4} = 1.065 * 10^{-3}$ M

$$K = \frac{[HI]^2}{[H_2]\,[I_2]} \rightarrow \frac{[9.35 * 10^{-4}]^2}{[6.5 * 10^{-5}]\,[1.065 * 10^{-3}]} = 51$$