CHE 1302

Basic Principles of Modern Chemistry II

Week 5

Hello and welcome to the weekly resources for Chemistry 1302! This resource covers topics typically taught by professors during the 5th week of classes.

On our website, https://baylor.edu/tutoring, you'll find the following links:

"Online Study Guide Resources" – If you don't see the topics you're learning right now, click here to find the weekly resources for the rest of the semester!

"How to Participate in Group Tutoring" - See if there is a Chemistry 1302 group tutoring session being hosted this semester – these are weekly question/answer sessions taught by our master tutors!

You can also view tutoring times for your course or schedule a private 30-minute appointment! Check out the website to learn more. You can also give us a call at (254)710-4135, or drop in. Our hours are Monday-Thursday 9 am – 8 pm on class days.
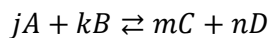
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KEY WORDS: The Equilibrium Condition & Constant, Le Chatelier's Principle

# TOPIC OF THE WEEK:

The Equilibrium Condition & Constant

Topic of the Week: The Equilibrium Condition & Constant

This week's resource is all about equilibrium! This topic applies to reversible reactions, such as:

$$jA + kB \rightleftarrows mC + nD$$

Equilibrium is the point at which the forward and reverse rates are equal.

1. A fair amount of reactant has been consumed; the rate of its consumption (forward rate) decreases because there is less of it to react.
2. A lot of product has been formed; the reverse rate increases because there is more of it to react.
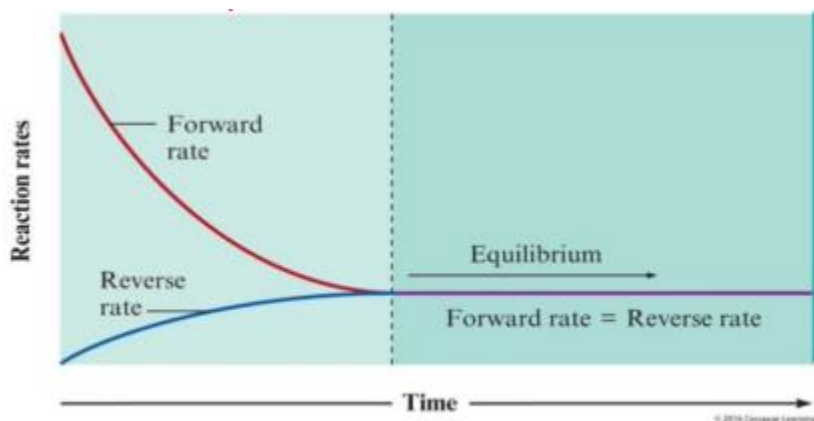
We need something that tells us *how far* an equilibrium reaction goes that will give us an idea of the relative amounts of reactants and products at equilibrium. This is where the equilibrium constant, K, comes in!

K is the ratio of products to reactants at equilibrium, calculated as: $K_c = \dfrac{[C]^m * [D]^n}{[A]^j * [B]^k}$

Note: K has no units

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Highlight 1: Calculating K in special cases

CASE 1: If you **have K for an equation**, and you are **trying to calculate K for a similar equation**, there are a couple of rules that can be followed:

1. If the new equation is the reverse of the first, $K_c' = 1/K_c$
   a. This is because, in your new calculation of the K equation, what used to be reactants would be products, and what used to be products would be reactants.
   b. Old Kc: $K_c = \dfrac{[C]^m * [D]^n}{[A]^j * [B]^k}$
   c. New Kc: $K_c' = \dfrac{A^j * [B]^k}{[C]^m * [D]^n}$
2. If the new equation is multiplied by r, $K_c' = K_c^r$
   a. Multiplying the new equation by r means that the coefficients will change:
   $$jrA + krB \rightleftarrows mrC + nrD$$
   b. So now, when calculating the Kc equation, each of the exponents are multiplied by r:
   $K_c' = \dfrac{[C]^{mr} * [D]^{nr}}{[A]^{jr} * [B]^{kr}}$
   c. This can be simplified to $K_c' = \left(\dfrac{[C]^m * [D]^n}{[A]^j * [B]^k}\right)^r = K_c^r$
3. Adding a couple of equations together: $K_c' = K_{c1} K_{c2}$
   a. Remember, K is calculated by multiplying concentrations of substrates together. Try writing the K expressions for a couple of different equations, adding the equations together, and writing the K expression for that total equation.

CASE 2: Equilibrium Expressions Involving Pressures:

For ==gases==, it can be more convenient to measure partial pressures than molarities.

So, the equilibrium expression for an equation involving gases can also be written as follows:

$$K_p = \frac{(P_G)^g * (P_H)^h}{(P_A)^a * (P_B)^b}$$

Also, $K_p = K_c * (RT)^{\Delta n}$. Some notes on this:

1. $\Delta n = (moles\ of\ gaseous\ products) - (moles\ of\ gaseous\ reactants)$
2. Check out the derivation for this equation, which uses PV=nRT:
   https://www.kpu.ca/sites/default/files/downloads/gasequil.pdf
3. Keep in mind that $K_p \neq K_c$ unless: RT=1 or $\triangle n = 0$

CASE 3: Heterogeneous equilibria (multiple phases present in one equation):

==Ignore solids and liquids== when calculating equilibrium expressions. Equilibrium can be defined by a certain ratio between the amount of reactant in solution and the amount of product in solution. Solids are not in solution; they are *not evenly distributed throughout* the liquid. And liquids are the *solvents in which the solutes are distributed* – they are also not said to be "in solution."

Do make sure to ==include these phases: (aq), (g)==

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Highlight 2: Disturbing K and Le Chatelier's

"==Q==" can be defined as the ratio of products to reactants at ==any point== (not just at equilibrium!)

$$Q = \frac{[C]^m * [D]^n}{[A]^j * [B]^k}$$

Le Chatelier's Principle: If the reaction is not at equilibrium, it will try to get there – and it will try to ==reverse any disturbance to equilibrium.== It "wants" to have a certain ratio of products to reactants, so if it doesn't have enough products, it will make more products, which means that it will move in the forward direction more than the reverse direction. And if it doesn't have enough reactants, it will make more of those, which means that it will move in the reverse direction more than the forward direction.

If:	Then:
Q=K	The reaction is at equilibrium
==Q<K==	There are more reactants than the reaction would like. So, ==the reaction will move in the net forward direction==/shift to the right. Product concentrations will increase, and reactant concentrations will decrease.
==Q>K==	There are more products than the reaction would like. So, ==the reaction will move in the net reverse direction==/shift to the left. Product concentrations

	will decrease, and reactant concentrations will increase.

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Highlight 3: Other Types of Disturbances

Changing concentration isn't the only thing that can shift equilibrium. Here are some other types of disturbances and their effects on the equilibrium.

Concentration changes (remember, addition of solids will not change anything!)	Q changes
[for gas systems]: pressure increases, or volume decreases	Equilibrium shifts toward the side of the equation that has fewer moles of gas. This "corrects" for the increase
Addition of an *inert* (doesn't react with anything) gas	Constant V: no change Constant P: pressures of reactants/products decrease
Change temperature: *Think of heat as one of the reactants/products. If the reaction is exothermic, heat is like a product. If the reaction is endothermic, heat is like a reactant. Now, remember Le Chatelier's principle.* Exothermic: $\quad reactants \leftrightharpoons products + heat$ Increasing temperature adds products. Endothermic: $\quad reactants + heat \leftrightharpoons products$ Increasing temperature adds reactants.	 Equilibrium shifts to the left. Equilibrium shifts to the right.
Addition of a catalyst	Lowers activation energy, increasing rate. Equilibrium is reached *faster*, but the balance between reactants and products (rate constant) stays the same. If the reaction is already at equilibrium, there is no change.

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Check Your Understanding

1. Write the expression for K$_c$: $2H_2S_{(g)} + CH_{4(g)} \rightleftharpoons CS_{2(g)} + 4H_{2(g)}$
2. At equilibrium at 525 K, there are 17.2 g PCl4, 16.4 g PCl3, and 8.84 g Cl2 in a 4.00 L container.
   a. What is the value of Kc for the reaction $PCl_5 \rightleftharpoons PCl_{3(g)} + Cl_{2(g)}$?

    b. What is the value of K_p?

3. Write the equilibrium constant expression for the reactions:

    a. $CaCO_{3(s)} \rightleftharpoons CaO_{3(s)} + CO_{2(g)}$

    b. $2NO_{(g)} \rightleftharpoons N_{2(s)} + O_{2(g)}$

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Things You May Struggle With

1. Don't forget that the concentrations you use to calculate K must be the concentrations **at equilibrium**. When calculating Q, this is not necessarily the case.
2. Check to make sure your chemical equation is balanced before using it to calculate K.
3. When you have a series of reactions for which you are manipulating K values to get an overall K value: If an equation is multiplied by a constant, K does not get multiplied by that same constant! Instead, it is raised to that power. (See Highlight 1, Case 1, #2)
4. Don't consider concentrations of solids and liquids when calculating an equilibrium expression – just use (aq) and (g).

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That's all this week! Please reach out if you have any questions and don't forget to visit the Tutoring Center website for further information at www.baylor.edu/tutoring. Answers to the Check Your Learning section are below.

1. $([H_2]^4[CS_3])/([H_2S])^2[CH_4])$

2.

    a. 0.0451

    b. 1.94

3.

    a. $[CO_2]$

    b. $\dfrac{O_2}{[NO_2]^2}$