5.2 - Equilibrium Constant - K_{eq}

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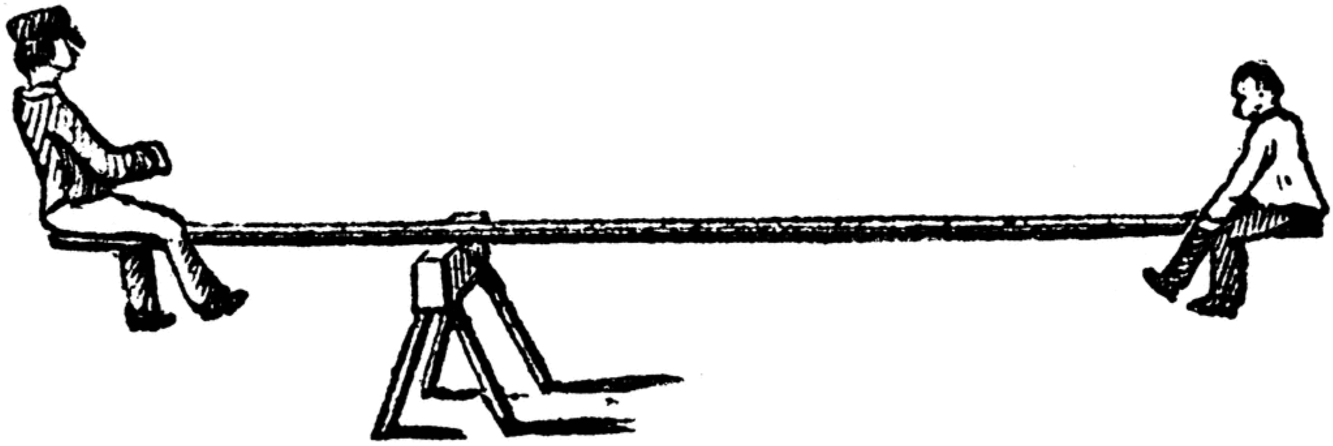


-Some chemical systems have basically no reaction, while others readily go to completion. However, most chemical systems fall somewhere in between these two extremes.

-Given the following chemical reaction; predict the number of moles of product produced:

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

The Law of Chemical Equilibrium

-states that a chemical system may reach a point in which a particular ratio of reactant and product concentrations has a constant value called the equilibrium constant (K_{eq} or K_c)

For a general reaction,

$aA + bB \rightleftarrows cC + dD$

where a,b,c,d are balancing coefficients and A,B,C,D are substances, an equilibrium constant (K_{eq}) expression can be written as:

$$K_{eq} = \frac{[C]^c \times [D]^d}{[A]^a \times [B]^b}$$

This mathematical relationship is true for all equilibrium systems.

Another term for this equation is the **mass-action expression**.

Ex) For the equilibrium reaction

$H_2 + I_2 \rightleftarrows 2\,HI$

find the equilibrium constant if $[H_2] = 0.022$ M, $[I_2] = 0.022$ M, and $[HI] = 0.156$ M.

Since K_{eq} is a constant for a reaction, it does not change unless the temperature of the system changes.

It does not matter on the initial concentrations used to reach equilibrium, just the concentrations **at equilibrium**.

For example, the following data was taken during an experiment with the equation H_2 + I_2 $\rightleftharpoons$ 2 HI at equilibrium:

Trial	[HI]	[H_2]	[I_2]	K_{eq}
1	0.156	0.0220	0.0220	50.3
2	0.750	0.106	0.106	50.1
3	1.00	0.820	0.0242	50.4
4	1.00	0.0242	0.820	50.4
5	1.56	0.220	0.220	50.3

Note that for this reaction, K_{eq} is always the same (ignoring experimental error) irrespective of the concentration of A, B, C, and D you started with.

When calculating K_{eq} for a given reaction, we do **not** include substances in the liquid or solid phase.

This is because the concentrations of substances in these phases do not change, but are constant no matter how much you have.

So, only include gaseous and aqueous states when calculating K_{eq}.

Homogeneous and Hetergeneous Equilibria Constants

-**Homogeneous Equilibrium** refers to a system with all components in the same phase

ex: $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

Write the equilibrium expression for the above reaction:

-**Hetegeneous Equilibrium** refers to a system with components that exist is more than one physical state

ex: $2NaHCO_{3(s)} \rightleftharpoons Na_2CO_{3(s)} + CO_{2(g)} + H_2O_{(g)}$

Write the equilibrium expression for the above reaction:

What does K_{eq} tell us?

There are 3 situations:

i. If K_{eq} is very large:

The concentration of the products are much greater than the concentration of the reactants. This means that the reaction essentially 'goes to completion'. That is, all - or most of - the reactants are used up to form the products.

Equilibrium lies to the right.

-We will call a number greater than 10^{10} very large.

For example he decomposition of ozone, O_3

$2\,O_{3(g)} \rightleftarrows 3\,O_{2(g)} \quad K_{eq} = 2.0 \times 10^{57}$

ii. If K_{eq} is very small:

The concentration of the products are much smaller than the concentration of the reactants. This means the reaction does not occur to a great extent. That is, most of the reactants remain unchanged because only a few products are formed.

Equilibrium lies to the left.

-We will call a value less than 10^{-10} very small.

For example, The production of nitrogen monoxide

$N_{2(g)} + O_{2(g)} \rightleftarrows 2\,NO_{(g)} \qquad K_{eq} = 1.0 \times 10^{-25}$

iii. If K_{eq} is neither very large or very small

This means that there significant amounts of both products and reactants formed at equilibrium.

We call values between 10^{-10} and 10^{10} neither very large or very small.

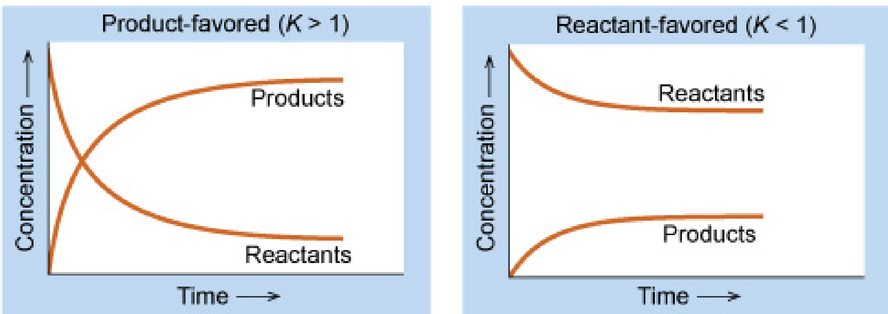
-K_{eq} >1: a bit more product at equilibrium

-K_{eq} >1: a bit more reactant at equilibrium

-K_{eq} =1: neither is favored

For example, the reaction of carbon monoxide and water

$$CO_{(g)} + H_2O_{(g)} \rightleftarrows CO_{2\,(g)} + H_{2\,(g)} \qquad K_{eq} = 5.09 \text{ (at 700 K)}$$



(a) (b)

5.2 – Equilibrium Constant – K_{eq} – Assignment

1. Write equilibrium expressions for the following reversible reactions:

a. $2\ NO_{2\ (g)} \leftrightarrow N_2O_{4\ (g)}$

b. $N_{2\ (g)} + 3\ H_{2\ (g)} \leftrightarrow 2\ NH_{3\ (g)}$

c. $H_2O_{(g)} + C_{(s)} \leftrightarrow H_{2(g)} + CO_{(g)}$

d. $2\ SO_{2\ (g)} + O_{2\ (g)} \leftrightarrow 2\ SO_{3\ (g)}$

e. $Cu_{(s)} + 2Ag^{+}{}_{(aq)} \leftrightarrow Cu^{2+}{}_{(aq)} + 2Ag_{(s)}$

2. For the equilibrium system described by $2\ SO_2$ (g) + O_2 (g) $\leftrightarrow$ $2\ SO_3$ (g) at a particular temperature the equilibrium concentrations of SO_2, O_2 and SO_3 were 0.75 M, 0.30 M, and 0.15 M, respectively. At the temperature of the equilibrium mixture, calculate the equilibrium constant, Keq, for the reaction.

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5.2 – Equilibrium Constant – K_{eq} – Assignment

3. For the equilibrium system described by: PCl_5 (g) ↔ PCl_3 (g) + Cl_2 (g) Keq equals 35 at 487°C. If the concentrations of the PCl_5 and PCl_3 are 0.015 M and 0.78 M, respectively, what is the concentration of the Cl_2?

4. Find the concentration of the products for the following:

$NH4Cl_{(s)} \leftrightarrow NH_{3(g)} + HCl_{(g)}$ when $K_{eq} = 6.0 \times 10^{-9}$

5. For the equilibrium reaction

$CO_{(g)} + H_2O_{(g)} \leftrightarrow CO_{2(g)} + H_{2(g)}$

the K_{eq} value at 690˚C = 10.0. A reaction mixture is analyzed and found to contain 0.80M CO, 0.050M H_2O, 0.50M CO_2, and 0.40M H_2. Show that the reaction is not at equilibrium.

6. For each of the following reactions, state whether the value of the equilibrium constant favours the formation of reactants or products.

a. $I_{2(g)} + Cl_{2(g)} \rightleftharpoons 2ICl_{(g)}$ $\quad K_{eq} = 2 \times 10^6$

b. $H_{2(g)} + Cl_{2(g)} \rightleftharpoons 2HCl_{(g)}$ $\quad K_{eq} = 1.08$

c. $I_{2(g)} \rightleftharpoons I_{(g)} + I_{(g)}$ $\quad K_{eq} = 3.8 \times 10^{-7}$