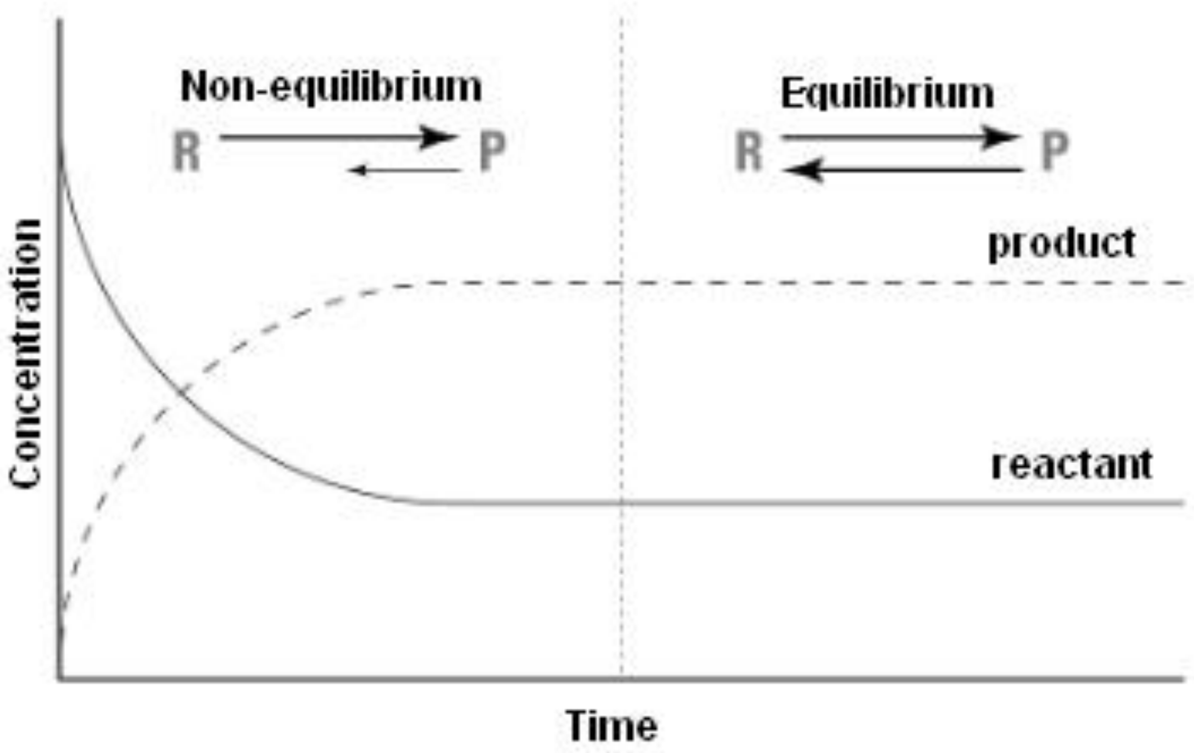
Equilibrium Reversible Reactions

Oak Park High School



Mrs. Kornelsen
Chemistry 40s

Intro: Do Blue bottle Reaction demo

Or watch online: http://www.dlt.ncssm.edu/core/Chapter14-Gas_Phase-Solubility-Complex_Ion_Equilibria/blue_bottle-lg.htm
Or
https://www.youtube.com/watch?v=kGSPAkOgN3U

Chemical & Physical Equilibrium

See PPT.

Equilibrium – A Dynamic State

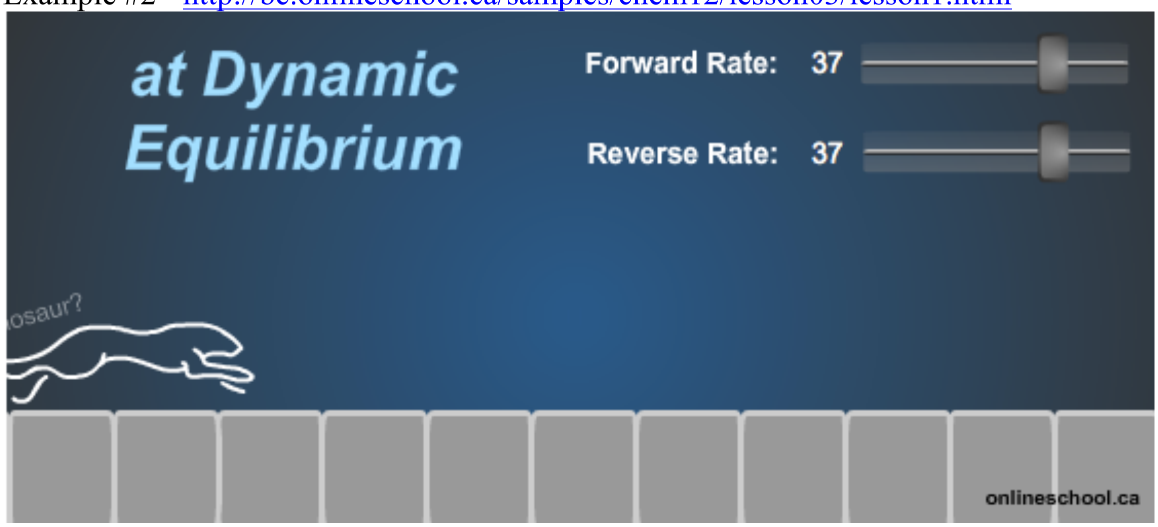
1. The difference between chemical and physical equilibrium.
    - Physical is a reversible change of state.
    - Chemical is a reversible chemical reaction that can go in the direction to create products or to create reactants.
2. Equilibrium can only happen in a **closed reaction**.
3. Equilibrium is only possible in **reversible reactions**. Chemists use a double arrow to represent these equations $\leftrightharpoons$.
4. Equilibrium is **dynamic** because the forward and reverse reactions are continuously happening at the **same time at the same rate**.
5. A reaction is at equilibrium once the forward and reverse reactions are happening **at the same rate**.
6. At equilibrium **concentrations remain constant**. This does not mean that the concentrations of all the reactants and the products are equal, just constant.

Equilibrium example - Imagine yourself on an escalator that is going down.

You start at the top (reactants) and end up at the bottom (products). But when you are partway down you start walking ***up*** the escalator as it continues going down. If you match your rate of walking ***up*** to the same rate that the escalator is going ***down***, you make no progress and appear to be at a standstill.

To an observer it would look as if you and the escalator had come to a stop, but actually both upward and downward movements continue. **Dynamic equilibrium!**

Example #2 - http://bc.onlineschool.ca/samples/chem12/lesson03/lesson1.html



Chemical Equilibrium

A chemical reaction is in equilibrium when **the quantities of reactants and products are no longer changing**.
The direction in which we write a chemical reaction (and thus which components are considered reactants and which are products) is not important. Thus the two equations

$H_2 + I_2 \rightarrow 2\,HI$ "synthesis of hydrogen iodide"

$2\,HI \rightarrow H_2 + I_2$ "dissociation of hydrogen iodide"

represent the same chemical reaction It makes no difference whether we start with two moles of HI or one mole each of H_2 and I_2; once the reaction has run to completion, the quantities of these two components will be the same.

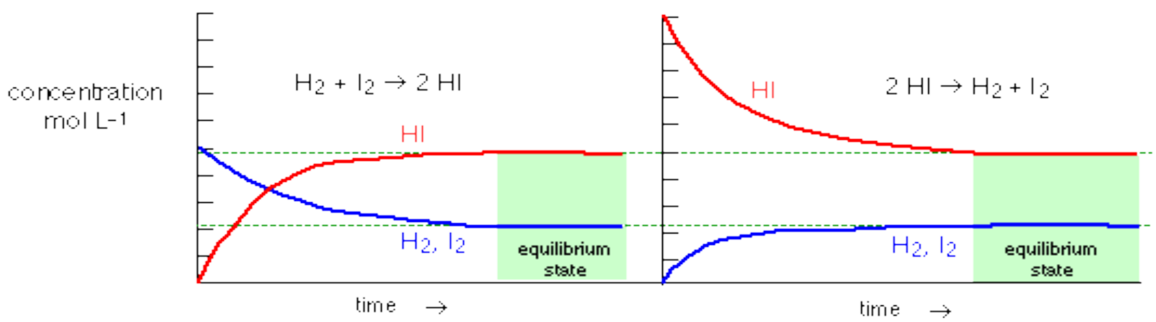
The two diagrams below show how the concentrations of the three components of this chemical reaction change with time.

In graph one which components start with a concentration of 0mol/L? _____________
Are these the products or reactants? ______________________________________

In graph two which components start with a concentration of 0mol/L? _____________
Are these the products or reactants? ______________________________________

This shows that these two graphs represent the same chemical reaction system, but with the reactions occurring in opposite directions.

Most importantly, note how the final (equilibrium) concentrations of the components are the same in the two cases.



The equilibrium state is independent of the direction from which it is approached. Whether we start with an mixture of H_2 and I_2 (left) or a pure sample of hydrogen iodide (shown on the right), the equilibrium state (shaded regions on the right) will be the same.

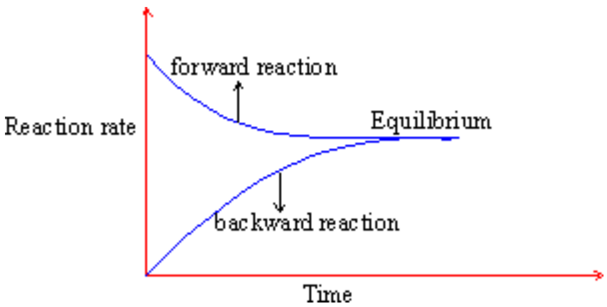
What is a reversible reaction?

A chemical equation of the form A → B represents the transformation of A into B.

A reversible chemical system or reaction is one that can proceed in **either direction**. So the **naming of products and reactants becomes subjective**; depending on which direction the reaction is proceeding. We will begin to use words like shifting left or shifting right to describe the direction of a reaction. Rather than saying the reaction is proceeding to the products side of the equation.

In a reversible reaction, the single arrow in the equation is replaced with a double arrow pointing in opposite directions, as in A $\rightleftharpoons$ B.

Let's practice:



1. Explain how you know this graph has reached equilibrium.

2. Which curve represents the products and which curve the reactants?

Class Activity:

Have a group of students represent sodium and chloride ions in the following reaction:
$NaCl$ + heat ←→ $Na^{+}_{(aq)} + Cl^{-}_{(aq)}$

For example, in a class of 20 students, 10 students could represent sodium ions and 10 students could represent chloride ions. Have 4 sodium ions and 4 chloride ions link arms on the left side of the room to represent sodium chloride particles. Have the remaining 12 students stand on the right side of the room. Ask a student to record on the board the number of each type of particles.

At this point, explain that in order for sodium chloride to break apart, heat is required. Place on the floor four pieces of red construction paper (to represent the heat), which can be picked up by the students representing the sodium chloride particles so that they can break up into sodium and chloride ions and move to the right side of the room. (The sodium ions in the sodium chloride particle should hold onto the heat). Students on the right side of the room could use the heat to join together to form a sodium chloride particle and move to the left side of the room.

Allow this movement to continue for a few minutes, and then have a student record the number of each particle a second time. Repeat this process once more so that students can see that equilibrium has occurred.
Emphasize that the process of equilibrium is not finished. The forward and reverse processes continue to occur

Practice:

1.

Compare:	Open and Closed Systems	
Contrast:	Open System	Closed System

2.

Compare:	Physical and Chemical Equilibrium	
Contrast:	Physical Equilibrium	Chemical Equilibrium

3. What is a reversible reaction?

4. What do the double arrows ($\rightleftharpoons$) mean in a chemical equation?

5. Explain why a rope-pull completition is similar to chemical equilibrium?

6. During an investigation, your lab partner states that a chemical system is not at equilibrium because the concentrations of the reactants and products are not equal. Explain why this statement might be incorrect.

7. The term *dynamic equilibrium* literally means "balanced change." Explain why this term describes a chemical system in equilibrium.

8. Carbon monoxide (CO) and hydrogen (H_2) react to form methane (CH_4) and water (H_2O). Write the chemical equation that represents this reversible reaction at equilibrium.

What is the Law of Mass Action?

The Law of Mass Action shows that **equilibrium can be approached from either direction** (see the hydrogen iodide illustration above), implying that any reaction $aA + bB \rightarrow cC + dD$ is really a **competition** between a "forward" and a "reverse" reaction.

Guldberg and Waage (don't need to know these names) showed that the rate of the reaction in either direction is **proportional to what they called the "active masses"** of the various components.

$$\text{rate of forward reaction} = k_f[A]^a[B]^b$$
$$\text{rate of reverse reaction} = k_r[C]^c[D]^d$$

In the equilibrium expression there is a proportionality constant *k* called the *rate constant* (remember kinetics) and quantities in square brackets representing the concentrations.

If we combine the two reactants A and B, the forward reaction starts immediately; then, as the products C and D begin to build up, **the reverse process gets underway**. As the reaction proceeds, the rate of the forward reaction diminishes while that of the reverse reaction increases. Eventually the two processes are **proceeding at the same rate**, and the reaction is at equilibrium:

$$\text{rate of forward reaction} = \text{rate of reverse reaction}$$
$$k_f[A]^a[B]^b = k_r[C]^c[D]^d$$

Again, remember that k is constant for a reaction, and can only be changed by temperature. Therefore $k_f = k_r$.

Both reactants and products will be present at any given point in time. The equilibrium may **favor either the reactants or products**. The extent to which the reaction proceeds towards products is measured by an equilibrium constant.

This constant is a specific ratio of the products to the reactants. This ratio is often referred to as a mass action expression.

$$aA + bB = cC + dD$$

where A and B are reactants
C and D are products
a,b,c,d are the coefficients in the balanced chemical equation

The mass action expression consists of the product of the products, each raised to the power given by the coefficient in the balanced chemical equation, over the product of the reactants, each raised to the power given by the coefficient in the balanced chemical equation. This mass action expression is set equal to the equilibrium constant, K_{eq}:

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

Let's practice:
Write the mass action expressions for the following reactions.

1. $2H_2 + O_2 \leftrightarrow 2H_2O$

2. $2NaCl + MgO \leftrightarrow Na_2O + MgCl_2$

3. $CH_3COOH + H_2O \leftrightarrow CH_3COO- + H_3O+$

4. $4W + 3X \leftrightarrow 5Y + 2Z$

Problem Solving Using Keq

For the reaction between hydrogen and iodine gas to produce hydrogen iodide:

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

The **equilibrium constant expression** will be:

$$K_{eq} = \frac{[HI]^2}{[H_2]\ [I_2]}$$

Equilibrium **concentrations** for each substance were measured at **equilibrium** and found to be:

At equilibrium:	$[H_2] = 0.022$ M $[I_2] = 0.022$ M $[HI] = 0.156$ M

We substitute these values into our equilibrium expression and solve for K_{eq}:

$$K_{eq} = \frac{[HI]^2}{[H_2]\ [I_2]} = \frac{(0.156)^2}{(0.022)(0.022)} = 50.3$$

The value of K_{eq}, which has no units, is a constant for any particular reaction, and its value does not change unless **the temperature of the system** is changed. This is the same as the rate constant k. It does not depend on the initial concentrations used to reach the point of equilibrium.For example, the following data were obtained for equilibrium concentrations of H_2, I_2 and HI; calculate the Keq of each trail.

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

Aside from accounting for slight experimental variation between trials, the value for $\mathbf{K_{eq}}$ **is the same** despite differences in equilibrium concentrations for the individual participants.

There is one other important point to make at this time.

- K_{eq} relates the concentrations of **products to reactants** at equilibrium.
- The concentrations of both aqueous solutions and gases change during the progress of a reaction. For reactions involving a solid or a liquid, while the ***amounts*** of the solid or liquid will change during a reaction, **their** *concentrations* (much like their densities) **will** *not* **change** during the reaction.
- Instead**, their values will remain constant**. Because they are constant, their values **are not included in the equilibrium constant expression**.

For example, consider the reaction showing the formation of solid calcium carbonate from solid calcium oxide and carbon dioxide gas:

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

The equilibrium constant for this reaction is (before changing):

$$K_{eq} = \frac{[CaCO_3]}{[CaO]\ [CO_2]}$$

But we remove those participants whose state is **either a solid or a liquid**, which leaves us with the following equilibrium constant expression:

$$K_{eq} = \frac{1}{[CO_2]}$$

Calculating K_{eq} Practice:

Calculate K_{eq} for each of the following. Be sure to set up the equilibrium constant expression first, before substituting in the values.

a) $H_2(g) + Cl_2(g) \rightleftarrows 2\ HCl$

$[H_2] = 1.0 \times 10^{-2}$ M

$[Cl_2] = 1.0 \times 10^{-2}$ M

$[HCl] = 1.0 \times 10^{-2}$ M

b) $N_2(g) + 3\ H_2(g) \rightleftarrows 2\ NH_3(g)$

$[N_2] = 4.4 \times 10^{-2}$ M

$[H_2] = 1.2 \times 10^{-1}$ M

$[NH_3] = 3.4 \times 10^{-3}$ M

c) $2\ CO(g) + O_2\ (g) \rightleftarrows 2\ CO_2\ (g)$

$[CO] = 2.5 \times 10^{-3}$ M

$[O_2] = 1.6 \times 10^{-3}$ M

$[CO_2] = 3.2 \times 10^{-2}$ M

d) $CH_4(g) + H_2O(g) \leftrightarrow CO(g) + 3\ H_2(g)$

$[CH_4] = 2.97 \times 10^{-3}$ M

$[H_2O] = 7.94 \times 10^{-3}$ M

$[CO] = 5.45 \times 10^{-3}$ M

$[H_2] = 2.1 \times 10^{-3}$ M

One more problem style to think about….

What is the concentration of O_2 if Keq = 5.0×10^{-21} and the concentrations of H_2 and H_2O are both 0.5M?

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

What is the concentration of HI if Keq = 6.7×10^{2} and the concentrations of H_2 and I_2 are both 0.25M and 0.7M respectively?

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

Equilibrium Constant (K_{eq}) Worksheet

Write the expression for the equilibrium constant K_{eq} for the reactions below in the left hand column and then solve for the equilibrium constant (K_{eq}).

K_{eq} Expression	Solve for K_{eq}
$N_{2(g)} + 3H_{2(g)} \leftrightarrow 2NH_{3(g)}$ $K_{eq} = \frac{[NH_3]^2}{[H_2]^3[N_2]}$	$[NH_3]$ = 0.01 M, $[N_2]$ = 0.02 M, $[H_2]$ = 0.02 M K_{eq} = 625
$2KClO_{3(s)} \leftrightarrow 2KCl_{(s)} + 3O_{2(g)}$ $K_{eq} = [O_2]^3$	$[O_2]$ = 0.05 M $[KCl]$ = 0.05 M $[KClO_3]$ = 0.01M $K_{eq} = 1.25 \times 10^{-4}$
$H_2O_{(l)} \leftrightarrow H^+_{(aq)} + OH^-_{(aq)}$ $K_{eq} = [OH^-][H^+]$	$[H^+] = 1 \times 10^{-8}$ M, $[OH^-] = 1. \times 10^{-6}$ M, $[H_2O] = 1. \times 10^{-6}$ M $K_{eq} = 1 \times 10^{-14}$
$2CO_{(g)} + O_{2(g)} \leftrightarrow 2CO_{2(g)}$ $K_{eq} = \frac{[CO]^2}{[O_2][CO]^2}$	$[CO]$ = 2.0 M, $[O_2]$ = 1.5 M, $[CO_2]$ = 3.0 M K_{eq} = 1.5
$Li_2CO_{3(s)} \rightarrow 2Li^+_{(aq)} + CO_3^{2-}{}_{(aq)}$ $K_{eq} = [CO_3^{2-}][Li^+]^2$	$[Li^+]$ = 0.2 M, $[CO_3^{2-}]$ = 0.1 M, $[Li_2CO_3]$ = 0.1 M $K_{eq} = 4 \times 10^{-3}$

The Meaning of K_{eq}

What can the value of K_{eq} tell us about a reaction?

- **If K_{eq} is very large**, the concentration of the products is much greater than the concentration of the reactants **(R<P)**. The reaction essentially "goes to completion"; all - or most of - of the **reactants are used up** to form the products.
- **If K_{eq} is very small**, the concentration of the reactants is much greater than the concentration of the products **(R>P)**. The reaction does not occur to any great extent - most of the reactants remain unchanged, and **there are few products produced**.
- **When K_{eq} is not very large or very small** (close to a value of 1) then roughly **equal amounts** of reactants and products are present at equilibrium.

Here are some examples to consider:

the decomposition of ozone, O_3	$2\ O_{3\ (g)} \rightleftarrows 3\ O_{2\ (g)}$ K_{eq} is very large, indicating that mostly O_2 is present in an equilibrium system, with very little O_3	$K_{eq} = 2.0 \times 10^{57}$
production of nitrogen monoxide	$N_{2\ (g)} + O_{2\ (g)} \rightleftarrows 2\ NO_{(g)}$ Very little NO is produced by this reaction; N_2 and O_2 do not react readily to produce NO (lucky for us - otherwise we would have little oxygen to breath in our atmosphere!)	$K_{eq} = 1.0 \times 10^{-25}$
reaction of carbon monoxide and water	$CO_{(g)} + H_2O_{(g)} \rightleftarrows CO_{2\ (g)} + H_{2\ (g)}$ The concentrations of the reactants are very close to the concentrations of the products at equilibrium	$K_{eq} = 5.09$ (at 700 K)

More Keq Problems

Quantitative Problems involving Keq

So far we have dealt with quantitative problems dealing with chemical equilibria. By examining the amounts of reactants and products at equilibrium, we can determine whether equilibrium favors either reactants or products.

Suppose we are given the following equilibrium at 500 K:

$$CO(g) + 2\ H_2\ (g) \rightleftharpoons CH_3OH(g)$$

The equilibrium concentrations are: $[CO] = 0.0911$ M, $[H_2] = 0.0822$ M, $[CH_3OH] = 0.00892$ M, what is the value of the equilibrium constant? Does the equilibrium favor reactants or products?

First, we need to write the mass action expression:

$$K_{eq} = \frac{[CH_3OH]}{[CO][H_2]^2}$$

Next, substitute the equilibrium concentrations into the mass action expression, and calculate for the equilibrium constant:

$$Keq = 14.5$$

Since the value of the equilibrium constant is **greater than one, (K_{eq} >1)**, the equilibrium favors the **products**.

Do the Equilibrium Constant Worksheets

Do p. 444 #31-36

Practice Questions

The Equilibrium Constant

1. Write balanced chemical equations for each of the following. Pay close attention to the physical states!

   Also - you must include the charge when writing ions, otherwise your answer is incorrect.

   Do not balance these equations using fractions for coefficients.

   sulfur dioxide gas combines with oxygen gas to produce sulfur trioxide gas

   carbon monoxide gas burns in gaseous oxygen to produce carbon dioxide gas

   hydrogen chloride gas is produced from hydrogen gas and chlorine gas.

   nitrogen gas and oxygen gas combine to produce gaseous dinitrogen oxide.

   solid hydrogen cyanide dissolves to produce hydrogen ions and cyanide ions in solution.

   solid silver chloride dissolves to produce silver ions and chloride ions in solution.

   calcium ions and phosphate ions come out of solution to produce solid calcium phosphate.

2. For each of the above reactions, write the equilibrium expression, Keq, for the reaction. Remember not to include solids or liquids in the equilibrium constant expression.

3. The equilibrium equation for the formation of ammonia is

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

   At 200°C the concentrations of nitrogen, hydrogen, and ammonia at equilibrium are measured and found to be:

   $[N_2] = 2.12$ $[H_2] = 1.75$ $[NH_3] = 84.3$

   Calculate K_{eq} at this temperature.

4. For each of the following equilibrium systems, identify whether the reactants or products are favored at equilibrium, or whether they are equally favored.

$COCl_{2\,(g)} \leftrightarrow CO_{(g)} + Cl_{2\,(g)}$	$K_{eq} = 8.2\times 10^{-2}$ at 627°C
$C_{(s)} + 2\ H_{2\,(g)} \leftrightarrow CH_{4\,(g)}$	$K_{eq} = 8.1\times 10^{8}$ at 25°
$PCl_{5\,(g)} \leftrightarrow PCl_{3\,(g)} + Cl_{2\,(g)}$	$K_{eq} = 2.24$ at 227°C
$H_{2\,(g)} + Cl_{2\,(g)} \leftrightarrow 2\ HCl_{(g)}$	$K_{eq} = 1.8 \times 10^{33}$ at 25°C
$C_{(s)} + H_2O_{\,(g)} \leftrightarrow CO_{(g)} + H_{2\,(g)}$	$K_{eq} = 1.96$ at 1000°C
$Mg(OH)_{2\,(s)} \leftrightarrow Mg^{2+}_{(aq)} + 2\ OH^{-}_{(aq)}$	$K_{eq} = 1.2\times 10^{-11}$ at 25°C

5. For the reaction: carbon monoxide burns in oxygen to produce carbon dioxide

   You are given the following equilibrium conditions:

   $[O_2] = 1.30 \times 10^{-3}$ $[CO_2] = 2.50 \times 10^{-4}$ $K_{eq} = 3.60 \times 10^{-3}$

   Calculate [CO]

Practice Questions Section 2.1

The Equilibrium Constant Answers

1. Write equilibrium expressions for the following reversible reactions:

Answers

a. $2\ NO_{2\ (g)} \leftrightarrow N_2O_{4\ (g)}$ $$K_{eq} = \frac{[N_2O_4]}{[NO_2]^2}$$

b. $N_{2\ (g)} + 3\ H_{2\ (g)} \leftrightarrow 2\ NH_{3\ (g)}$ $$K_{eq} = \frac{[NH_3]^2}{[N_2]\ [H_2]^3}$$

c. $2\ SO_{2\ (g)} + O_{2\ (g)} \leftrightarrow 2\ SO_{3\ (g)}$ $$K_{eq} = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

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, K_{eq}, for the reaction.

Solution:

Begin by setting up the equilibrium constant expression for the balanced equation:

$$K_{eq} = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

Next, substitute in the known values, and solve for the unknown, which is Keq for this question. Don't forget to use the exponents!

$$K_{eq} = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{(0.15)^2}{(0.75)^2(0.30)} = 0.13$$

3. For the equilibrium system described by: $PCl_{5\,(g)} \leftrightarrow PCl_{3\,(g)} + Cl_{2\,(g)}$

   K_{eq} equals 35 at 487°C. If the concentrations of the PCl5 and PCl3 are 0.015 M and 0.78 M, respectively, what is the concentration of the Cl_2?

   Solution:

   Again begin by setting up the equilibrium constant for the equation and then substitute in the know values. But in this case, K_{eq} is one of our known values; one of the concentrations is the unknown.

$$K_{eq} = \frac{[PCl_3]\ [Cl_2]}{[PCl_5]}$$

   Let χ = the unknown, [Cl_2]. Substitute in known values and solve for χ:

$$35 = \frac{(0.78)\ (\chi)}{(0.015)}$$

$$35 \times 0.015 = 0.78 \times (\chi)$$

$$\chi = [Cl_2] = 0.67M$$

4. The following table gives some values for reactant and product equilibrium concentrations (in moles/L; M) at 700 K for the Shift reaction, an important method for the commercial production of hydrogen gas:

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

   Calculate K_{eq} for each of the five trials. How do the answers compare with each other? Why?
   Solution:

   Set up the equilibrium constant expression and solve for Keq for each trial. $K_{eq} = \dfrac{[CO_2]\ [H_2]}{[CO]\ [H_2O]}$

Trial	[CO_2]	[H_2]	[CO]	[H_2O]	K_{eq}
1	0.600	0.600	0.266	0.266	5.09
2	0.600	0.800	0.330	0.286	5.09
3	2.00	2.00	0.877	0.877	5.09
4	1.00	1.50	0.450	0.655	5.09
5	1.80	2.00	0.590	1.20	5.09

Keq = 5.09 for all trials.

Keq is a constant and will not change unless the temperature of the system is changed (there are often minor variations resulting from experimental measurements, however).

The Reaction Quotient Q *and the Equilibrium Constant* K_{eq}

If the system is NOT at equilibrium, the ratio you set up does not equal the equilibrium constant. In such cases, the ratio is called a **reaction quotient** which is designated as *Q*.

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

The closer a reaction gets to achieving equilibrium the closer the value of *Q* will become to the value of the equilibrium constant, *K*

$Q \rightarrow K_{eq.}$

The Reaction Quotient "Q" is the Numerical Value of the Equilibrium expression at any condition (not necessarily at equilibrium). You can calculate Q from the concentrations and or pressures given. Then compare it with K_{eq}. Remember a system is trying to achieve equilibrium.

- If **$K<Q$** then the reaction will shift to make more reactants, (to the left) to establish equilibrium again. **(If Q is bigger, we have passed equilibrium and need to shift back to the left)**
- If **$K>Q$** then the reaction will shift to make more products, (to the right) to establish equilibrium again. **(If K is bigger, we have not yet achieved equilibrium and need to shift forward or right)**
- If **$K=Q$** then the system is at equilibrium already.

Hint - Write K on the left, Q on the right, and then put the $<$ or $>$ in, If it points to the left, the reaction stays on the left, if it points to the right, the reaction tends to the right.

The expression for the reaction quotient, Q, looks like that used to calculate equilibrium constants but Q can be calculated for any set of conditions.

In order to determine Q we need to know:

- The equation for the reaction, including the physical states.
- The molarities of the different species all measured at the same moment in time.

To calculate Q:

- Write the expression for the reaction quotient, same as the expression for Keq.
- Find the concentrations of each species involved.
- Substitute values into the expression and solve.

Example: 0.035 moles of SO_2, 0.500 moles of SO_2Cl_2, and 0.080 moles of Cl_2 are combined in an evacuated 5.00 L flask and heated to 100°C. What is Q before the reaction begins? Which direction will the reaction proceed in order to establish equilibrium?

$$SO_2Cl_2(g) \leftrightarrows SO_2(g) + Cl_2(g) \qquad K_{eq} = 0.078 \text{ at } 100^oC$$

- Write the expression to find the reaction quotient, Q.

$$Q_c = \frac{[SO_2][Cl_2]}{[SO_2Cl_2]}$$

- The amounts must be expressed as moles per liter, so a conversion is required.

0.500 mole SO_2Cl_2/5.00 L = 0.100 M SO_2Cl_2
0.035 mole SO_2/5.00 L = 0.0070 M SO_2
0.080 mole Cl_2/5.00 L = 0.016 M Cl_2

- Substitute the values in to the expression and solve for Q.

Q = 0.0011

- Compare the answer to the value for the equilibrium constant and predict the shift.

0.078 (K) > 0.0011 (Q)

Since K >Q, the reaction will proceed in the forward direction in order to increase the concentrations of both SO_2 and Cl_2 and decrease that of SO_2Cl_2 until Q = K.

Practice Questions:

1. For the reaction below, K is 1.8 at a certain temperature. If the initial concentrations of all species are 0.50 M, which direction will the reaction move in?

$2NO(g) + O_2(g) < = > 2NO_2(g)$

2. At 300 °C, $K_{eq} = 8.50$ for the all gas-phase reaction $SiO + H_2O \quad <=> \quad SiO_2 + H_2$

Initially 0.80 moles of each of the four species are mixed together in a two litre flask. Calculate the equilibrium concentrations of all four species calculate Q and determine if the reaction is at equilibrium. If not which direction does the system need to shift to achieve equilibrium?

3. An equilibrium system was kept at constant temperature and pressure in a **ten litre** container. It contained 0.100 mole of SO_2 , 0.200 mole of NO_2 , 0.300 mole of NO, and 0.500 mole of SO_3. What is the equilibrium constant for the following reaction?

$SO_3\ (g) \ + \ NO\ (g) \quad <=> \quad SO_2\ (g) \ + \ NO_2\ (g)$

If Q = 200 is the reaction at equilibrium? If not which direction does the system need to shift to achieve equilibrium?

4. In the following all-gaseous reaction at equilibrium, the concentration of each substance is [A] = 2 M; [B] = 3 M; [C] = 3 M; [D] = 1 M. What is the value of the equilibrium constant for the following reaction?

$2\,A \ + \ 2\,B \ <=> \ 3\,C \ + \ 2\,D$

If Q = 1.0 is the reaction at equilibrium? If not which direction does the system need to shift to achieve equilibrium?

5. For the reaction below, K is 107 at a certain temperature. If the initial concentrations of all species are 0.10 M, which direction will the reaction move in?

$$2H_2O(l) <=> O_2(g) + 2H_2(g)$$

6. At 927°C Kc = 3.91 for the reaction: $CO(g) + 3\ H_2(g) \rightleftharpoons CH_4(g) + H_2O(g)$
Initially $[CO] = [H_2] = 0.0200$ M and $[CH_4] = [H_2O] = 0.00100$ M.
Under these initial conditions, is this reaction at equilibrium? If not, which direction will the reaction proceed to attain equilibrium?

7. At 1000 °C, Keq = 1.17 for the reaction: $CO_2(g) + C(s) \rightleftharpoons 2\ CO(g)$
A 10.0 L vessel contains 1.5x10-3 mol of CO_2 and 5.5x10-3 mol of CO. Is the system at equilibrium?

8. Consider the reaction: $2\ NOBr(g) \rightleftharpoons 2\ NO(g) + Br_2(g)$
Keq = 3.07×10^{-4} at 24°C. For each of the following sets of initial conditions, describe in which direction the reaction will proceed to reach equilibrium.

Set 1	Set 2	Set 3
[NOBr] = 0.0610 M	[NOBr] = 0.115 M	[NOBr] = 0.500 M
[NO] = 0.0151 M	[NO] = 0.0169 M	[NO] = 0.0170 M
$[Br_2]$ = 0.0108 M	$[Br_2]$ = 0.0142 M	$[Br_2]$ = 0.0140 M

9. A mixture of 1.57 mol of N_2, 1.92 mol of H_2, and 8.13 mol of NH_3 is introduced into a 20.0 L reaction vessel at 500 K. At this temperature, the equilibrium constant for the reaction $N_2(g) + 3\ H_2(g) \rightleftharpoons 2NH_3\ (g)$ is 1.7×10^2. Is the reaction mixture at equilibrium? If not, what is the direction of the net reaction?

ICE Charts

This is the next big thing in chemistry…ICE charts!!!!!

Problem type 1

When the initial concentrations of some species, usually **reactants are given** and the **equilibrium concentration of one of the products** is given, calculate the **equilibrium constant**.

For the reaction $H_2(g) + F_2(g) \leftrightarrow 2HF(g)$

1.00 moles of hydrogen and 1.00 moles of fluorine are sealed in a 1.00 L flask at 150°C and allowed to react. At equilibrium, 1.32 moles of HF are present. Calculate the equilibrium constant.

Note - we need to put **concentration values** into our ICE chart.

Solution:

To solve this problem, it is best to set up a table recording **initial concentrations** (**I**), **change** in concentrations (**C**) and **equilibrium** concentrations (**E**), **ICE** for short.

Since the flask is 1.0 L, $[H_2] = [F_2] = 1.0$ mol/L and the initial concentration of HF is **zero**.

We set up the table by first rewriting the equation and ICE down the left side. We also insert the initial concentrations.

	$H_2(g)$	$+ F_2(g)$	$2HF(g)$
I	1.00	1.00	0.00
C			
E			

In our **change row** we plug in + or – x to represent whether that species is decreasing or increasing.

We can not simply plug in x we must plug in x multiplied by its **corresponding coefficient**. I told you that stoichiometry was coming back.

Above we would plug in –x, -x, +2x

	$H_2(g)$	$+ F_2(g)$	$2HF(g)$
I	1.00	1.00	0.00
C	-x	-x	+2x
E			

The equilibrium value for HF is given and we can now plug it in.

	$H_2(g)$	$+ F_2(g)$	$2HF(g)$
I	1.00	1.00	0.00
C	-x	-x	+2x
E			1.32

Now we can solve for x and **find the equilibrium concentrations** of the reactants.

2x = 1.32

x = .66

To calculate the equilibrium concentrations of the reactants we need to **subtract x from our initial concentrations**.

$1.00\text{-}x = [H_2]_{equilibrium}$ $\qquad$ $1.00\text{-}x = [F_2]_{equilibrium}$

$1.00\text{-}0.66 = [H_2]_{equilibrium}$ $\qquad$ $1.00\text{-}0.66 = [F_2]_{equilibrium}$

This means the equilibrium concentrations of H_2 and F_2 are **0.34 mol/L** and HF is 1.32 mol/L

Fill these last values into our ICE chart.

	$H_2(g)$	$+ F_2(g)$	$2HF(g)$
I	1.00	1.00	0.00
C	-x	-x	+2x
E	0.34	0.34	1.32

Next, we substitute the equilibrium concentrations into the equilibrium law and solve for K_{eq}:

$$K_c = \frac{[HF]^2}{[H_2][F_2]} = \frac{(1.32)^2}{(0.34)(0.34)} = \frac{1.742}{0.1156}$$

$$K_c = 15.1$$

Let's do #1 and #14 from the exercises on pg 25 together. You may want to do these on separate paper for extra room.

Do practice questions # 3, 4, 7, 14 & 15, they are examples of problem type 1. Next class we will do the rest which represent problem type 2.

Problem type 2

Ok now we can make it a little tougher….

Given Initial Concentrations and K, Find Equilibrium Concentrations

If we are given initial concentrations and the equilibrium constant, we can **calculate the equilibrium concentrations** of all reactants and products for many equilibrium systems.

For the reaction below the equilibrium constant is 6.76. If 6.0 moles of nitrogen and oxygen gases are placed in a 1.0 L container, what are the concentrations of all reactants and products at equilibrium?

$$N_2(g) + O_2(g) \leftrightarrow 2\ NO(g)$$

Solution:

Step 1. Set up and ICE table.

	$N_2(g)$	$+ O_2(g)$	$2\ NO(g)$
I	6.0 M	6.0 M	0.00 M
C			
E			

Here we do not know any of the equilibrium concentrations. However, we do know from stoichiometry, as the reaction proceeds **equal amounts of nitrogen and oxygen** will be consumed and **twice** that amount of **NO will be produced**. Therefore we can fill in our **change row**. Since we do not know the amounts consumed, we will assign them a value of x and the amount of NO produced 2x.

	$N_2(g)$	$+ O_2(g)$	$2\ NO(g)$
I	6.0	6.0	0.00
C	$-x$	$-x$	$+2x$
E			

Complete the equilibrium row as well including our variables.

	$N_2(g)$	$+ O_2(g)$	$2\ NO(g)$
I	6.0	6.0	0.00
C	$-x$	$-x$	$+2x$
E	$6.0 - x$	$6.0 - x$	$2x$

Step 2. Write the equilibrium law and substitute equilibrium values from the ICE table.

We will use the **equilibrium values** that we filled into our equilibrium row of (6.0 – x) and (2x). Note Kc is how some chemistry disciplines refer to Keq.

$$K_c = \frac{[NO]^2}{[N_2][O_2]}$$

$$6.76 = \frac{(2x)^2}{(6.0 - x)(6.0 - x)}$$

Step 3. Solve for x to determine equilibrium concentrations.

The denominator is a square, so we can take the square root of both sides to avoid a quadratic equation.

$$\sqrt{6.76} = \sqrt{\frac{(2x)^2}{(6.0 - x)^2}}$$

$$2.60 = \frac{2x}{6.0 - x}$$

Solve for x by multiplying both sides by the denominator. Then isolate the x's

$$(6.0 - x)(2.60) = \left(\frac{2x}{6.0 - x}\right)(6.0 - x)$$

$$15.6 - 2.60x = 2x$$

$$15.6 = 2x + 2.60x$$

$$\frac{15.6}{4.6} = \frac{4.6x}{4.6}$$

$$3.4 = x$$

Step 4. Find Equilibrium concentrations.

Since x equals **the loss in concentrations** of nitrogen and oxygen,

Equilibrium $[N_2] = [O_2] = 6.0 – x = 6.0 – 3.4 = 2.6$ mol/L

Equilibrium [NO] = 2x = 2(3.4 mol/L) = 6.8 mol/L

Therefore, at equilibrium nitrogen and oxygen are both 2.6 mol/L and the concentration of NO is 6.8 mol/L.

For this course questions of this type (in this unit) will always be perfect squares so we do not have to use quadratic equations.

Questions 10-13 in Exercises are prime examples of these questions.

Practice Problem:

Suppose you are given the following equilibrium:

$$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g) \qquad K_{eq} = 23.2$$

If the initial amounts of CO and H_2O were both 0.100 M, what will be the amounts of each reactant and product at equilibrium?

	CO (g)	+ H_2O ⇌	CO_2 (g)	+ H_2(g)
Initial	0.100 M	0.100 M	0	0
Change				
Equilibrium				

Complete the chart:

	CO (g)	+ H_2O ⇌	CO_2 (g)	+ H_2(g)
Initial	0.100 M	0.100 M	0	0
Change	- x	-x	x	x
Equilibrium	0.100 -x	0.100 -x	x	x

Substitute the above algebraic quantities into the mass action expression and solve for x:

$$K_{eq} = \frac{[CO_2][H_2]}{[CO][H_2O]} = \frac{(x)(x)}{(0.100 - x)(0.100 - x)} = \frac{x^2}{(0.100 - x)^2} = 23.2$$

$$\sqrt{\frac{x^2}{(0.100 - x)^2}} = \sqrt{23.2}$$

$$\frac{x}{0.100 - x} = 4.85$$

Multiply both sides by the denominator, 0.100 - x:

$$\cancel{(0.100 - x)} \times \frac{x}{\cancel{(0.100 - x)}} = 4.85(0.100 - x)$$

The term, 0.100 - x, cancels out on the left hand side, and the term, 4.85 must be distributed through the term 0.100 - x on the right hand side:

$$x = 0.485 - 4.85x$$

Add 4.85x to both sides of the equation:

$$x + 4.85x = 0.485 - 4.85x + 4.85\,x$$

Combining terms:

$$5.85x = 0.485$$

Solve for x by dividing both sides by 5.85:

$$x = \frac{0.485}{5.85} = 0.0829 = [CO_2] = [H_2]$$

Recall that x represents the equilibrium quantities of both H_2 and CO_2 . The equilibrium quantities of CO and H_2O is given by:

$$0.100 - x = 0.100 - 0.0829 = 0.017\text{ M} = [CO] = [H_2O]$$

Equilibrium Calculations Worksheet

Complete the following questions on a piece of loose leaf. See work in teacher booklet

1. For the reaction: $H_{2(g)} + CO_{2(g)} \leftrightarrows H_2O_{(g)} + CO_{(g)}$

   Initially 1.00 mol of H_2 and 1.00 mol CO_2 were placed in a sealed 5.00 L vessel. What would be the equilibrium concentrations if K_{eq} is 0.772?

[H_2] and [CO_2] = 0.1065M [H_2O] and [CO] = 0.0935 M

2. A 10.0 L flask is filled with 0.200 mol of HI at 698 K. What will be the concentration of H_2, I_2 and HI at equilibrium? The value of K_{eq} is 0.0184.

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

[H_2] and [I_2] = 0.0192 M [HI] = 8.23×10^{-4} M

3. Phosphorus pentachloride, $PCl_{5(g)}$, dissociates at high temperature into phosphorous trichloride, $PCl_{3(g)}$, and chlorine, $Cl_{2(g)}$. Initially, 0.200 mol of PCl_5 were placed in a 5.00 L flask at 200°C, and at equilibrium the concentration of PCl_5 was found to be 0.015 M. Calculate the value of the equilibrium constant at 200°C.

K_{eq} = 0.0416

4. The equilibrium constant at 490°C for the reaction $2\,HI_{(g)} \leftrightarrows H_{2(g)} + I_{2(g)}$ K_{eq} is 0.022. What are the equilibrium concentrations of HI, H_2, and I_2, when an initial amount of 2.00 mol of HI gas is placed in a 4.3 L flask at 490°C.

[HI] = 0.3272M [H_2] and [I_2] = 0.069 M

5. Calculate K_{eq} for the decomposition of H_2S from this reaction:

$$2\,H_2S_{(g)} \leftrightarrows 2\,H_{2(g)} + S_{2(g)}$$

A tank initially contained H_2S with a concentration of 10.00 mol/L at 800 K. When the reaction had come to equilibrium, the concentration of S_2 gas was 2.0×10^{-2} mol/L.

6. The equilibrium constant, K_{eq}, for the reaction: $N_2O_{4(g)} \leftrightarrows 2\ NO_{2(g)}$ at 25°C is 5.88×10^{-3}. Suppose 15.6 g of N_2O_4 is placed in a 5.00 L flask at 25°C.
   Calculate:
   a) The number of moles of NO_2 present at equilibrium.
   b) The percentage of the original N_2O_4 that is dissociated.

Recall: $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$

Exercises – Predicting Equilibrium

Answer the following questions. Be sure to show all your work.

1. For the reaction below, K = 25.0:

   $A(s) + 2\ B(g) \leftrightarrow 2\ C(g)$

   If [A] = 12.0 mol/L [B] = 2.0 mol/L [C] = 30.0 mol/L

   a) Is the system is at equilibrium.
   b) Which reaction is faster (favoured), forward or reverse?
   c) Which concentrations are increasing or decreasing?

2. There exists an equilibrium if 5.0 moles of CO_2, 5.0 moles of CO and 0.20 moles of O_2 are in a 2.0 L container at 562°C. Find K_C for the reaction

   $2\ CO(g) + O_2(g) \leftrightarrow 2\ CO_2(g)$

   Would the system be at equilibrium if $[CO_2]$ = 15.8 mol/L, [CO] = 10.0 mol/L and $[O_2]$ = 0.25 mol/L? If not, which reaction is favoured?

3. For the reaction below, K = 16.0:

   $2\ SO_2(g) + O_2(g) \leftrightarrow 2\ SO_3(g)$

   Initially, $[SO_2]$ = 5.0 mol/L, $[O_2]$ = 10.0 mol/L and $[SO_3]$ = 0. After two hours $[O_2]$ = 7.9 mol/L. Is the system at equilibrium? If not, which substances are increasing and which are decreasing?

4. The reaction

   $4\ NH_3(g) + 5\ O_2(g) \leftrightarrow 4\ NO_2(g) + 6\ H_2O(g)$

   is at equilibrium when $[H_2O]$ = 0.100 mol/L, $[O_2]$ = 2.00 mol/L, [NO] = 0.200 mol/L and $[NH_3]$ = 0.500 mol/L. If 0.75 moles of H_2O, 12.0 moles of NO, 30.0 moles of O_2 and 0.30 moles of NH_3 are in a 3.0 L container at the same temperature, is equilibrium achieved? If not, which reaction is favoured?

5. K = 46.0 for the reaction

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

   Initially there are 6.90 moles of H_2 and 2.40 moles of I_2 in a 1.00 L container. After 5 hours there is still 1.00 moles of I_2 left. Is the system at equilibrium? If not, which substances are increasing and which are decreasing?

Exercises

1. A mixture at equilibrium at 827°C contains 0.552 moles of CO_2, 0.552 moles H_2, 0.448 moles CO, and 0.448 moles of H_2O in a 1.00 L container. What is the value of the equilibrium constant, K_{eq}?

$$CO_2(g) + H_2(g) \leftrightarrow CO(g) + H_2O(g)$$

$K_{eq} = 0.6587$

2. The equilibrium constant for the reaction

$$4\ H_2(g) + CS_2(g) \leftrightarrow CH_4(g) + 2\ H_2S(g)$$

at 755°C is 0.256. What is the equilibrium concentration of H_2S if at equilibrium $[CH_4]$ = 0.00108 mol/L, $[H_2]$ = 0.316 mol/L, $[CS_2]$ = 0.0898 mol/L?

$[H_2S] = 0.461$ M

3. Find the value of K if at equilibrium there is 25.0 moles of P_4, 10.0 moles of H_2 and 5.00 moles of PH_3, in a 5.00 L container. The equation is

$$P_4(g) + 6\ H_2(g) \leftrightarrow 4\ PH_3(g)$$

$K_{eq} = 0.003125$

4. Find the value of K for the equilibrium system

$$ZnO(s) + CO(g) \leftrightarrow Zn(s) + CO_2(g)$$

if at equilibrium there are 3.0 moles of CO, 4.0 moles of Zn and 4.0 moles of CO_2 in a 500.0 mL container.

$K_{eq} = 1.3333$

5. If K = 46.0 for

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

what $[I_2]$ would be in equilibrium with 0.50 mol/L HI and 0.10 mol/L H_2?

Only do if covered Q

6. If K = 10.0 for

$$N_2(g) + 3\ H_2(g) \leftrightarrow 2\ NH_3(g)$$

how many moles of NH_3, at equilibrium, will be in a 2.00 L container if $[H_2]$ is 0.600 mol/L and $[N_2]$ is 0.100 mol/L?

$[NH_3] = 0.4648$ M $NH_3 = 0.9295$ mol

7. The formation of ammonia from hydrogen and nitrogen occurs by the reaction below:

$$3\ H_2(g) + N_2(g) \leftrightarrow 2\ NH_3(g)$$

Analysis of an equilibrium mixture of nitrogen, hydrogen, and ammonia contained in a 1.0 L flask at 300°C gives the following results: hydrogen 0.15 moles; nitrogen 0.25 moles: ammonia 0.10 moles. Calculate K_{eq} for the reaction.

$K_{eq} = 11.85$

~~8. Bromine chloride, BrCl, decomposes to form bromine and chlorine.~~

$$2\ BrCl(g) \leftrightarrow Cl_2(g) + Br_2(g)$$

At a certain temperature the equilibrium constant for the reaction is 11.1, and the equilibrium mixture contains 4.00 mol of Cl_2 in a 1 Litre flask. How many moles of Br_2 and BrCl are present in the equilibrium mixture?

~~9. The decomposition of hydrogen iodide to hydrogen and iodine occurs by the reaction~~

$$\cancel{2\ HI(g) \leftrightarrow H_2(g) + I_2(g)}$$

~~Hydrogen iodide is placed in a container at 450°C an equilibrium mixture contains 0.50 moles of hydrogen iodide. The equilibrium constant is 0.020 for the reaction. How many moles of iodine and hydrogen iodide are present in the equilibrium mixture?~~

10. $H_2(g) + Cl_2(g) \leftrightarrow 2\ HCl(g)$

A student places 2.00 mol H_2 and 2.00 mol Cl_2 into a 0.500 L container and the reaction is allowed to go to equilibrium at 516°C. If K_C is 76.0, what are the equilibrium concentrations of H_2, Cl_2 and HCl?

$[H_2]$ and $[Cl_2]$ = 0.747M   [HCl] = 6.5 M

11. If K = 78.0 for the reaction

$$A(s) + 2\ B(g) \leftrightarrow 2\ C(g)$$

and initially there are 5.00 moles of A and 4.84 moles of B in a 2.00 L container, how many moles of B are left at equilibrium?

[C] = 2.17 M   [B] = 0.246 M

12. For the reaction:

    $C(s) + O_2(g) \leftrightarrow CO_2(g)$ K = 25.0

    Find the moles of CO_2 at equilibrium, if initially there are 100.0 moles of C, 50.0 moles of O_2 and 2.0 moles of CO_2 in a 2.00 L container.

    $[O_2]$ = 1M $[CO_2]$ = 24 M

13. For the reaction:

    $NH_4Cl(s) \leftrightarrow NH_3(g) + HCl(g)$ K = 3.50 x 10-4

    Find the concentration of NH_3 in a 1.00 L container at equilibrium if initially there were 0.200 moles of NH_3 added to 0.200 moles of HCl.

    $[NH_3]$ and [HCl] = 0.0187 M

14. Initially the concentrations of N_2 and O_2 are 1.8 mol/L each and there is no NO. If at equilibrium the [NO] is 2.0 mol/L, find K.

    $$N_2(g) + O_2(g) \leftrightarrow 2\ NO(g)$$

    $[N_2]$ and $[O_2]$ = 0.8M [NO] = 2.0 M

15. Find K for the reaction

    $$2\ CO(g) + O_2(g) \leftrightarrow 2\ CO_2(g)$$

    if initially, there is 5.0 moles of CO, 10.0 moles of O_2 and 1.0 mole of CO_2 in a 2.0 L container and at equilibrium CO_2 has a concentration of 2.5 mol/L.

    [CO] = 0.5 M $[O_2]$ = 4 M $[CO_2]$ = 2.5 M

A Whole Mix of K_{eq} Problems

1. If K = 25.0 for the reaction and initially there are 5 moles of A and 8 moles of B in a 1.00 L container, what is the concentration of B at equilibrium?

$$A(s) + 2\,B(g) \leftrightarrow 2\,C(g)$$

2. For the reaction:

$C(s) + O_2(g) \leftrightarrow CO_2(g)$ K = 25.0

Find the concentration of CO_2 at equilibrium, if initially there are 100.0 moles of C and O_2 and 2.0 moles of CO_2 in a 2.00 L container.

3. For the reaction:

$NH_4Cl(s) \leftrightarrow NH_3(g) + HCl(g)$ K = 1.25

Find the concentration of NH_3 at equilibrium if initially the concentration of NH_3 and HCl is 0.5M.

4. Initially the concentrations of N_2 and O_2 are 2.25 mol/L each and there is no NO. If at equilibrium the [NO] is 4.0 mol/L, find K.

$$N_2(g) + O_2(g) \leftrightarrow 2\,NO(g)$$

5. Find K for the reaction

$$2\,CO(g) + O_2(g) \leftrightarrow 2\,CO_2(g)$$

if initially, there is 15.0 moles of CO, 15.0 moles of O_2 and no CO_2 in a 2.0 L container. At equilibrium CO_2 has a concentration of 5 mol/L.

Do page 451 #51-59
Page 444 #31-36
Page 461 #81-84

Review of Equilibrium Part A: Multiple Choice

1. Dynamic Equilibrium prerequisites include

a) a closed system b) at least a little bit of all species present at eqm
c) rate of forward reaction equals rate of reverse reaction
d) reaction is reversible e) all of these

2. The equilibrium constant

a) is a positive value b) changes with temperature
c) is calculated from the mass-action expression
d) has gas and aqueous phase terms put in it e) all of the above

3. Which of the following stresses cannot create a new set of equilibrium concentrations?

a) pressure b) volume c) concentration d) temperature e) catalyst

4. Which one of the following statements is **true**?

a. reactant and product concentrations are constant with time in a system at chemical equilibrium.
b. equilibrium is dynamic, with two opposing chemical reactions occurring at different rates.
c. the nature and properties of the equilibrium state depend on the direction from which the equilibrium is reached.
d. a system at chemical equilibrium reacts so as to maximize slight disturbances.

5. An equilibrium system was kept at constant temperature and pressure in a **five litre** container. It contained 0.500 mole of SO_2 , 0.500 mole of NO_2 , 0.250 mole of NO, and 0.500 mole of SO_3. What is the equilibrium constant for the following reaction:

$$SO_3\,(g) + NO\,(g) \Leftrightarrow SO_2\,(g) + NO_2\,(g)$$

a. 4.0 b. 2.3 c. 0 d. 0.25 e. – 3.14

6. Which of the following K expressions correctly describes the behaviour of a saturated silver carbonate (Ag_2CO_3) solution in water? $Ag_2CO_{3(s)} \rightarrow 2Ag^+_{(aq)} + CO_3^{2-}{}_{(aq)}$

a. $K = [Ag+][CO_3^{2-}]$ b. $K = [Ag+]\,[CO_3^{2-}]^2$ c. $K = [Ag^+]^2\,[CO_3^{2-}]$

d. $K = [Ag+][CO_3^{2-}] / [Ag_2CO_3]$ e. $K = [Ag^+]^2\,[CO_3^{2-}]/ [Ag_2CO_3]$

7. If Keq=4 for the following system which set of equilibrium concentrations could be valid?

$$3\,A\,(g) + 2\,B\,(aq) \Leftrightarrow C(l) + 5\,D\,(g)$$

a. A=2, D=1, B=1 b. A=1, D=1, B=1 c. A=2, B=1, D=2

Chemical Equilibrium Part 1 Review – more multiple choice

1. Consider the chemical reaction A(g) + B(g) → AB(g) + heat; K = 0.50
As the concentration of B is increased at constant temperature, the value of
K will 1) decrease; 2) increase; 3) remain the same; 4) none of these.

2. Which of the following statements concerning chemical equilibrium is **incorrect**?
   1) Chemical equilibrium can occur at different temperatures.
   2) Chemical equilibrium may be established quickly.
   3) Chemical equilibrium may be established slowly.
   4) When chemical equilibrium is established, the reaction stops.

3. A chemical reaction has reached equilibrium when
   1) the reverse reaction begins;
   2) the forward reaction ceases;
   3) the concentrations of the reactants and products become equal;
   4) the concentrations of the reactants and products become fixed.

4. In a reversible reaction, if the forward reaction is nearly complete before the speed of the reverse reaction becomes high enough to establish equilibrium,
   1) the products of the forward reaction are favored.
   2) the original reactants are favored.
   3) considerable concentrations of both reactants and products are present.
   4) the new products are favored.

5. For reversible reactions at equilibrium, a rise in temperature will
   1) favor the endothermic reaction;
   2) decrease the speed of the reaction;
   3) not affect the value of the equilibrium constant;
   4) result in the liberation of heat.

6. What is the effect of a catalyst on the reaction $N_2(g) + 3\ H_2(g) \rightleftarrows 2\ NH_3(g)$?
   1) It produces more product.
   2) It changes the time needed to reach equilibrium.

REVIEW WRITTEN RESPONSES

1. Write the K_{eq} expression for each of the following equilibrium systems:

a) $SrCO_3$ (s) <=> SrO (s) + CO_2 (g)

b) 2 Cu^+ (aq) + CO_3^{2-} (aq) <=> Cu_2CO_3 (s)

c) H_2O (g) + F_2 (s) <=> 2 HF (g) + O (g)

2. At 639 kelvin the all gas-phase system 2 SeO_2 + O_2 <=> 2 SeO_3 (g) is examined in a **three litre** flask. It is found that at equilibrium there are 0.60 moles of SeO_2 , 0.90 moles of O_2 and 0.70 moles of SeO_3 . Calculate K_{eq}

3. At 900 kelvin, consider the all gas-phase reaction

2 A + B <=> 3 C + D

Initially, 0.60 M A and 0.60 M B are mixed together. (No C or D is present). When equilibrium is eventually reached, the equilibrium concentration of D is found to be 0.10 M. Calculate K_{eq} .

4. At 300 °C, K_c = 8.50 for the all gas-phase reaction

SiO + H_2O <=> SiO_2 + H_2

Initially 0.80 moles of each of the four species are mixed together in a two litre flask.

Calculate the equilibrium concentrations of all four species and the percent reaction of silicon monoxide.