Chemistry 30 Unit 4 Acid Base / Equilibrium

The Equilibrium Law

Review: Le Chatelier's Principle

The following equilibrium takes place in a rigid container:

$$PCl_{5(g)} + \text{heat} \rightleftharpoons PCl_{3(g)} + Cl_{2(g)}$$

In what direction does the reaction shift as a result of:

a) adding $PCl_{5(g)}$? Ans: [redacted]
b) removing $Cl_{2(g)}$? Ans: [redacted]
c) decreasing temperature? Ans: [redacted]
d) increasing pressure? Ans: [redacted]
e) adding a catalyst? Ans: [redacted]

The Effect of Catalysts on Equilibrium

A catalyst is a compound added to a reaction which makes the reaction occur more quickly or at a lower activation energy.

The catalyst does not effect the equilibrium. It speeds up the forwards and reverse reactions equally, so equilibrium is not effected.

The Law of Chemical Equilibrium

Last day, we studied how stresses can effect a system in dynamic equilibrium.

We looked at the qualitative changes to the system (i.e. which way the reaction went), but we didn't consider any quantitative changes (i.e. how much the concentrations were changing by).

Today, we will study the changes in concentration of a equilibrium system using the Law of Equilibrium.

The principles which dictate concentrations in equilibrium were first developed by brothers-in-law Cato Guldberg and Peter Waage at the end of the 19th century.

They found that in a chemical system at equilibrium, there is a constant ratio between the [products] and the [reactants].

Consider the general reaction at equilibrium below:

$$aA + bB \rightleftharpoons cC + dD$$

where:

A and B = reactants
C and D = products
a, b, c and d = numerical coefficients

If these variables are arranged as a ratio, we get the Equilibrium Law expression:

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

where: K = the equilibrium constant for this particular system

Using this equation, if we know the concentrations of the products and reactants, we can find the equilibrium constant for the system.

This can then be used to solve for other missing concentrations later on.

ex) Set up the Law of Equilibrium expression for the following reactions:

a) $2\ SO_{2(g)} + O_{2(g)} \rightleftharpoons 2\ SO_{3(g)}$

b) $4\ NH_{3(g)} + 5\ O_{2(g)} \rightleftharpoons 4\ NO_{(g)} + 6\ H_2O_{(g)}$

ex) A mixture of nitrogen and chlorine gas is kept in a 5.0 L flask:

$$N_{2(g)} + 3Cl_{2(g)} \rightleftharpoons 2NCl_{3(g)}$$

When equilibrium is reached, the flask was found to contain 0.0070 mol N_2, 0.0022 mol Cl_2 and 0.95 mol NCl_3. Find the equilibrium constant for this equation.

Step 1: Find the [ ] of each reactant.

Step 2: Set up the equilibrium equation.

***Note: the units here DO NOT cancel out! Unit analysis does not apply!**

Ans: [redacted]

***The equilibrium constant is a unit less number.**

ex) Nitrogen and hydrogen gases were mixed in a 3500 mL flask to produce ammonia gas. At equilibrium, the mixture contained 0.25 mol of ammonia and 0.080 mol of hydrogen. The $K_c = 5.81 \times 10^5$. What number of moles of nitrogen was present at equilibrium?

Ans: [redacted]

Evaluating the K

Sometimes, the constant is given a subscript, such as

K_{eq} = simple equilibrium constant
K_c = concentration equilibrium
K_a = acid equilibrium
K_w = water equilibrium

The size of the equilibrium constant can also give information about the reaction:

$K_{eq} > 1$ - products favored: more products will be produced.
$K_{eq} < 1$ - reactants are favored: more reactants will be produced.
$K_{eq} = 1$ - neither favored.

Important Note

Concentrations of Pure Liquids and Solids

The idea of [ ] really only applies to aqueous solutions and gases. Solids and pure liquids (i.e. $Hg_{(l)}$) can not be expressed using [ ] as they are not dissolved in water.

When writing the equation, you can either omit pure liquids or solids or set their [ ] = 1 (which will not affect the equation).

Practice: WS