Acid/Base and Chemical Equilibrium

College of Health and Medical Techniques

Department of Anesthesiology

2nd Grade

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6074

Lecture 6

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Acids and Bases

In 1923, J. N. Brønsted in Denmark and J. M. Lowry in England proposed independently a theory of acid/base behavior that is especially useful in analytical chemistry.

According to the Brønsted-Lowry theory, **an acid is a proton donor, and a base is a proton acceptor.**

For a molecule to behave as an acid, it must encounter a **proton acceptor** (or base). Likewise, a molecule that can accept a proton behaves as a base if it encounters an **proton donor**

Conjugate Acids and Bases

An important feature of the Brønsted-Lowry concept is the idea that **the product formed when an acid gives up a proton is a potential proton acceptor** and is called the **conjugate base** of the parent acid. For example, when the species acid1 gives up a proton, the species base1 is formed, as shown by the reaction

acid1 = **base1** + proton

We refer to acid1 and base1 as a **conjugate acid/base pair**, or just a **conjugate pair**.

Similarly, every base accepts a proton to produce a **conjugate acid**. That is,

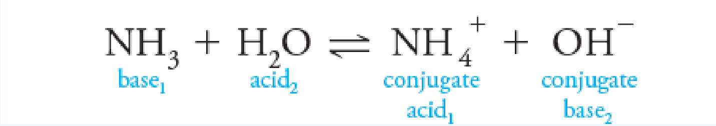
base2 + proton = **acid2**

When these two processes are combined, the result is an acid/base, or **neutralization**, reaction

acid1 + base2 = base1 + acid2

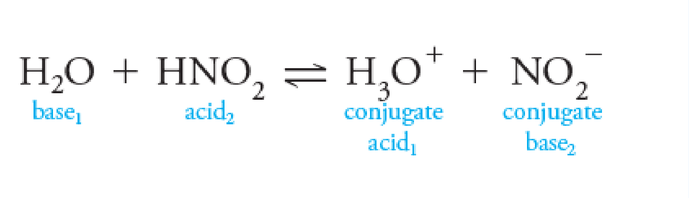
Conjugate Acids and Bases

Many solvents are proton donors or proton acceptors and can thus induce basic or acidic behavior in solutes dissolved in them. For example, in an aqueous solution of ammonia, water can donate a proton and acts as an acid with respect to the solute NH3



$$\underset{\text{base}_1}{NH_3} + \underset{\text{acid}_2}{H_2O} \rightleftharpoons \underset{\text{conjugate acid}_1}{NH_4^+} + \underset{\text{conjugate base}_2}{OH^-}$$

In this reaction, ammonia (base1) reacts with water, which is labeled acid2, to give the conjugate acid ammonium ion (acid1) and hydroxide ion, which is the conjugate base (base2) of the acid water. On the other hand, water acts as a proton acceptor, or base, in an aqueous solution of nitrous acid:



$$\underset{\text{base}_1}{H_2O} + \underset{\text{acid}_2}{HNO_2} \rightleftharpoons \underset{\text{conjugate acid}_1}{H_3O^+} + \underset{\text{conjugate base}_2}{NO_2^-}$$

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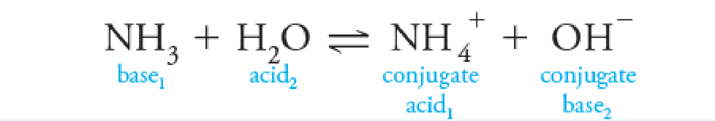

$$\underset{\text{base}_1}{NH_3} + \underset{\text{acid}_2}{H_2O} \rightleftharpoons \underset{\substack{\text{conjugate}\\ \text{acid}_1}}{NH_4^+} + \underset{\substack{\text{conjugate}\\ \text{base}_2}}{OH^-}$$

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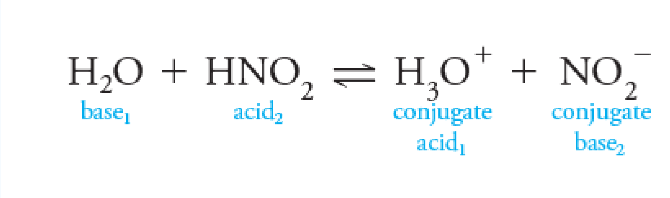
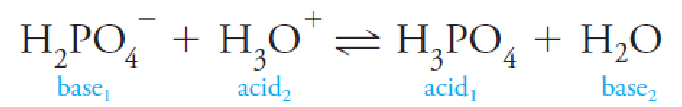


$$\underset{\text{base}_1}{H_2O} + \underset{\text{acid}_2}{HNO_2} \rightleftharpoons \underset{\substack{\text{conjugate}\\ \text{acid}_1}}{H_3O^+} + \underset{\substack{\text{conjugate}\\ \text{base}_2}}{NO_2^-}$$

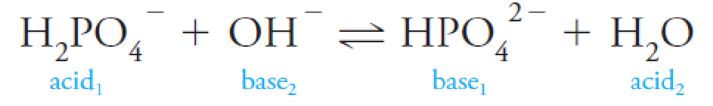
Amphiprotic Species

Amphiprotic Species are <u>Species that have both acidic and basic properties are amphiprotic</u>.

An example is dihydrogen phosphate ion,

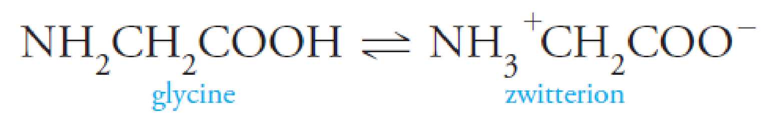


$$\underset{\text{base}_1}{H_2PO_4^-} + \underset{\text{acid}_2}{H_3O^+} \rightleftharpoons \underset{\text{acid}_1}{H_3PO_4} + \underset{\text{base}_2}{H_2O}$$



$$\underset{\text{acid}_1}{H_2PO_4^-} + \underset{\text{base}_2}{OH^-} \rightleftharpoons \underset{\text{base}_1}{HPO_4^{2-}} + \underset{\text{acid}_2}{H_2O}$$

The simple **amino acids** are an important class of **amphiprotic compounds that contain both a weak acid and a weak base functional group**. When dissolved in water, an amino acid, such as glycine, **undergoes a kind of internal acid/base reaction** to produce a **zwitterion**—a species that has both a positive and a negative charge. Thus.

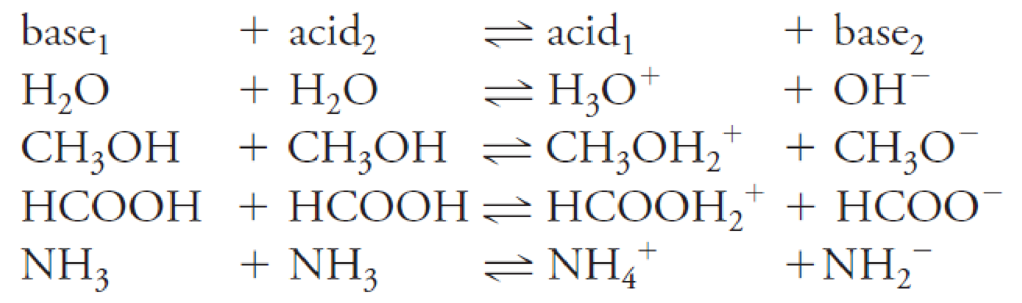


$$\underset{\text{glycine}}{NH_2CH_2COOH} \rightleftharpoons \underset{\text{zwitterion}}{NH_3^+CH_2COO^-}$$

Autoprotolysis

Amphiprotic solvents undergo self-ionization, or **autoprotolysis**, to form a pair of ionic species. Autoprotolysis is yet another example of acid/base behavior, as illustrated by the following equations

Autoprotolysis (also called autoionization) is the spontaneous reaction of molecules of a substance to give a pair of ions



base_1	$+ \text{acid}_2$	$\rightleftharpoons \text{acid}_1$	$+ \text{base}_2$
H_2O	$+ H_2O$	$\rightleftharpoons H_3O^+$	$+ OH^-$
CH_3OH	$+ CH_3OH$	$\rightleftharpoons CH_3OH_2^+$	$+ CH_3O^-$
$HCOOH$	$+ HCOOH$	$\rightleftharpoons HCOOH_2^+$	$+ HCOO^-$
NH_3	$+ NH_3$	$\rightleftharpoons NH_4^+$	$+ NH_2^-$

Chemical Equilibrium

Suppose the following reaction

$$aA + bB\ cC + dD$$

In this hypothetical reaction, substance A reacts with substance B (forward reaction) at a certain rate equal to...K1[A]a[B]b to produce substances C and D, but at the same time the resulting substances C and D return to the reaction (back or reverse reaction) to re-produce the reactants at a rate equal to K2[C]c[D]d. With the passage of time, the concentrations decrease. The reacting materials, and as a result, the speed of their reaction decreases, the concentrations of the resulting materials increase, and the speed of their reaction increases until the equilibrium point, at which the speed of the forward and reverse reactions is equal, and then the concentrations of the reactants and products remain constant.

Chemical Equilibrium

$$\boldsymbol{Rate\ 1 = K_1[\mathrm{A}]^a[\mathrm{B}]^b}$$

$$\boldsymbol{Rate\ 2 = K_2[\mathrm{C}]^c[\mathrm{D}]^d}$$

$$\boldsymbol{Rate\ 1 = Rate\ 2\ in\ Equilibrium}$$

$$\boldsymbol{K_1[\mathrm{A}]^a[\mathrm{B}]^b = K_2[\mathrm{C}]^c[\mathrm{D}]^d}$$

Rearranging the equation results

$$\boldsymbol{\frac{K_1}{K_2} = \frac{[\mathrm{C}]^c[\mathrm{D}]^d}{[\mathrm{A}]^a[\mathrm{B}]^b} = K_{eq}}$$

<u>Keq is the equilibrium constant and is a constant value no matter how the concentrations of the reactants or products change.</u>

<u>The equilibrium constant is affected only if the temperature or pressure changes</u>

Chemical Equilibrium

When the pressure increases, the system tries to get rid of the pressure according to Le Chatelier's principle, so the reaction moves towards the direction that reduces the number of moles of the substances, i.e. reduces the volume, and vice versa. As a result, the equilibrium position moves to a new position.

When the temperature increases, the reaction moves towards the direction that leads to heat absorption and vice versa.

The effect of temperature change on the equilibrium position and the value of the equilibrium constant show in table in the next page

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)\ \Delta H = -570\ \text{kJ}$$

Chemical Equilibrium

A positive ΔH means an endothermic reaction

Status	Direction of displacement of the reaction	Value of the equilibrium constant
The increase in temperature (heating)	towards the front or right	increases
The temperature decreases (cooling)	towards the left or back	decreases

Negative ΔH means exothermic reaction

The temperature decreases (cooling)	towards the front or right	increases
The increase in temperature (heating)	towards the left or back	decreases

The number of moles of reactants is greater than the number of moles of products

Status	Direction of displacement of the reaction	Value of the equilibrium constant
Increased pressure (decreased volume)	towards the front or right	increases
Decreased pressure (Increased volume)	towards the left or back	decreases

The number of moles of reactants is smaller than the number of moles of products

Decreased pressure (Increased volume)	towards the front or right	increases
Increased pressure (decreased volume)	towards the left or back	decreases

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