

## Brønsted-Lowry Theory of Acids and Bases

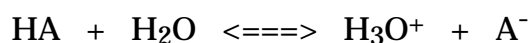
Acids and bases are substances which upset the  $K_w$  balance of  $[H^+]$  and  $[OH^-]$  in water:

$$K_w = [H_3O^+][OH^-]$$

Acids will increase the amount of  $H^+$  present, while bases increase the amount of  $OH^-$  present. As one of the two values ( $[H^+]$  and  $[OH^-]$ ) changes in one direction, the other will change in the opposite direction, so as to preserve the value of the constant  $K_w$ .

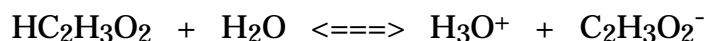
This theory defines an acid as a substance that will donate a proton to another substance. The receiving substance, called proton acceptor, is the base.

In general, for all acids in this theory, the following equation applies:



Note that, in the forward direction of this reaction, HA is an acid because it donates a proton to  $H_2O$ , which does the receiving. Therefore,  $H_2O$  is acting in the role of base.

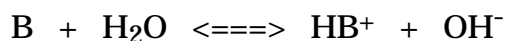
To use a specific example, let A be the acetate ion. The above reaction for acetic acid is:



However, it is important to also consider the reverse equation, since this is an equilibrium. In this case, note that  $H_3O^+$  is the acid, since it has a proton to donate.  $A^-$  is the base, since it will receive the proton.

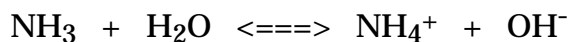
$H_3O^+$  is called the conjugate acid of  $H_2O$ , while  $A^-$  is called the conjugate base of HA.

In general, for all bases in this theory, the following equation applies:



Note that, in the forward direction of this reaction,  $H_2O$  is an acid because it donates a proton to B, which does the receiving. Therefore, B is acting in the role of base.

To use a specific example, let B be ammonia. The above reaction for ammonia is:



However, it is important to also consider the reverse equation, since this is an equilibrium. In this case, note that  $HB^+$  is the acid, since it has a proton to donate.  $OH^-$  is the base, since it will receive the proton.

$HB^+$  is called the conjugate acid of B, while  $OH^-$  is called the conjugate base of  $H_2O$ .

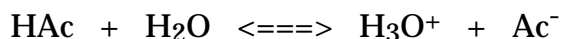
An acid (or a base) and its conjugate will differ in formula by only one proton. The acid of the two will have the proton, while the base of the conjugate pair lacks the proton.

## Treatment of Weak Acids in Brønsted-Lowry Terms

Acetic acid will be used as the example for this discussion. Almost all monoprotic weak acids can be treated in the following manner. Exceptions to the treatment below include very dilute solutions and extremely weak acids. Both of these situations require the use of a more exact equation which is beyond the scope of the current discussion.

The formula for acetic acid is  $\text{CH}_3\text{COOH}$  (sometimes also written as  $\text{HC}_2\text{H}_3\text{O}_2$ ). In the first formula, the acidic hydrogen is the last one while in the second formula, the acidic hydrogen is written first. This discussion will use an abbreviation for acetic acid, HAc. After the acid ionizes,  $\text{Ac}^-$  will symbolize the acetate ion (the formula for which is  $\text{CH}_3\text{COO}^-$  or  $\text{C}_2\text{H}_3\text{O}_2^-$ ).

When HAc ionizes in water, the following equation can be written:



An equilibrium constant can be written for this reaction, as follows:

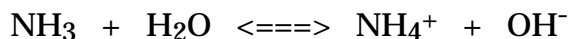
$$K_c = \frac{[\text{H}_3\text{O}^+][\text{Ac}^-]}{[\text{HAc}][\text{H}_2\text{O}]}$$

Since the value for  $[\text{H}_2\text{O}]$  is a constant, move it to the left side of the equation. Since  $K_c [\text{H}_2\text{O}]$  is itself a constant, we rename the new constant as  $K_a$ . We now have:

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{Ac}^-]}{[\text{HAc}]}$$

The  $K_a$  values for many acids have been measured. The  $K_a$  for acetic acid is  $1.76 \times 10^{-5}$ .

In like fashion, a  $K_b$  expression can be written for weak bases. When ammonia,  $\text{NH}_3$ , ionizes in water, the following equation can be written:



An equilibrium constant can be written for this reaction, as follows:

$$K_c = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3][\text{H}_2\text{O}]}$$

Since the value for  $[\text{H}_2\text{O}]$  is a constant, move it to the left side of the equation. Since  $K_c [\text{H}_2\text{O}]$  is itself a constant, we rename the new constant as  $K_b$ . We now have:

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

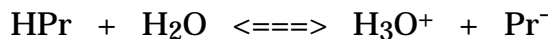
The  $K_b$  values for many bases have been measured. The  $K_b$  for ammonia is  $1.78 \times 10^{-5}$ .

## Writing $K_a$ and $K_b$ Expressions

I. Write the  $K_a$  expression for the following weak acids:

Acid Name	Abbreviation to use	$K_a$ value
Anisic	HAnis	$3.38 \times 10^{-5}$
Benzoic	HBz	$6.46 \times 10^{-6}$
Chloroacetic	HClac	$1.40 \times 10^{-3}$
Citric	HCi	$7.10 \times 10^{-4}$ (first step only)
Formic	HFor	$1.77 \times 10^{-4}$
Octanoic	HOc	$1.28 \times 10^{-5}$
Propionic	HPr	$1.23 \times 10^{-5}$
Theobromine	HThe	$1.29 \times 10^{-8}$
Thioacetic	HThi	$4.68 \times 10^{-4}$
Uric	HUr	$1.29 \times 10^{-4}$

In each case, the letters following the H stand for the anion, just like in the acetic acid example used previously. For example, propionic acid dissociates into  $H_3O^+$  and  $Pr^-$ , in this fashion:

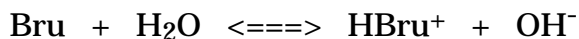


By the way, this is not the  $K_a$  expression for propionic acid! Nor is the example below the  $K_b$  expression for brucine!

II. Write the  $K_b$  expression for the following weak bases:

Acid Name	Abbreviation to use	$K_b$ value
Aniline	Anil	$2.34 \times 10^{-5}$
Brucine	Bru	$5.24 \times 10^{-9}$
Codeine	Cod	$6.15 \times 10^{-9}$
L-Ephedrine	Eph	$1.10 \times 10^{-10}$
Imidazol	Imd	$1.11 \times 10^{-7}$
Morphine	Mor	$6.16 \times 10^{-9}$
Papaverine	Pap	$3.98 \times 10^{-7}$
Purine	Pur	$5.01 \times 10^{-3}$ (first step only)
Quinine	Qui	$7.41 \times 10^{-5}$ (first step only)
Strychnine	Str	$5.49 \times 10^{-9}$

In each case, the letters stand for the unionized base. For example, Brucine ionizes into  $HBru^+$  and  $OH^-$ , in this fashion:



III. Compute the pH for a 0.100 M solution of each weak acid listed in I. above.

IV. Compute the pH for a 0.100 M solution of each weak base listed in II. above. Keep in mind you have to calculate the  $[OH^-]$  first, so there is an additional step needed to get the pH.