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6-1 The Reaction of a Bronsted–Lowry Acid with a Bronsted–Lowry Base

When a Bronsted–Lowry acid reacts with a Bronsted–Lowry base, a proton is transferred from the acid to the base.

The Bronsted–Lowry acid donates a proton to the Bronsted–Lowry base, which accepts it.

Consider, for example, the reaction of the general acid H A with the general base B: In an acid–base reaction, one bond is broken and one bond is formed. The electron pair of the base B: forms a new bond to the proton of the acid, forming H B$^+$

The acid(H A) loses a proton, leaving the electron pair in the H$^-$ A bond on A, forming A:

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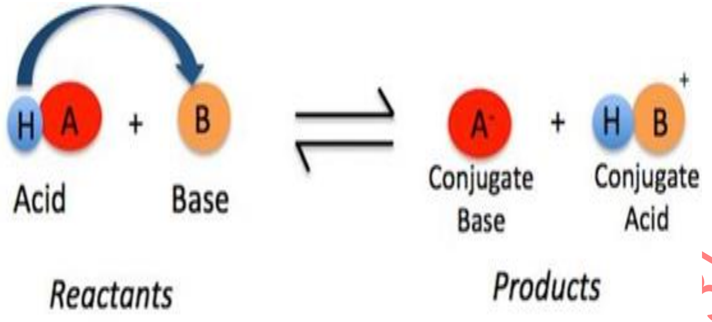

The product formed by loss of a proton from an acid is called its conjugate base.

The product formed by gain of a proton by a base is called its conjugate acid.

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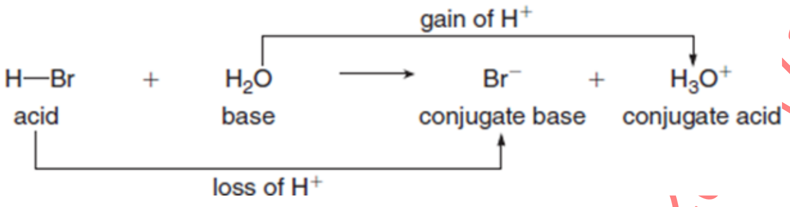
Thus, the conjugate base of the acid HA is A^-

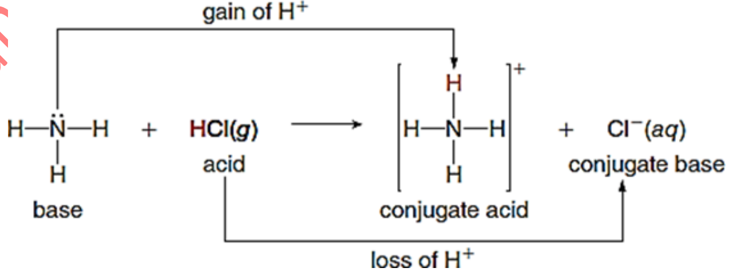
The conjugate acid of the base B: is HB^+



The reaction of ammonia (NH_3) with HCl is also a Bronsted–Lowry acid–base reaction.

In this example, NH_3 is the base since it gains a proton to form its conjugate acid, NH_4^+.

HCl is the acid since it donates a proton, forming its conjugate base, Cl^-.



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6-2 Acid and Base Strength.

Although all Bronsted–Lowry acids contain protons, some acids readily donate protons while others do not.

Similarly, some Bronsted–Lowry bases accept a proton much more readily than others.

How readily proton transfer occurs is determined by the strength of the acid and base.

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When a covalent acid dissolves in water, proton transfer forms (H_3O^+) and an anion.

The splitting apart of a covalent molecule (or an ionic compound) into individual ions is called dissociation.

Acids differ in their tendency to donate a proton; that is, acids differ in the extent to which they dissociate in water.

A strong acid readily donates a proton.

When a strong acid dissolves in water, essentially 100% of the acid dissociates into ions.

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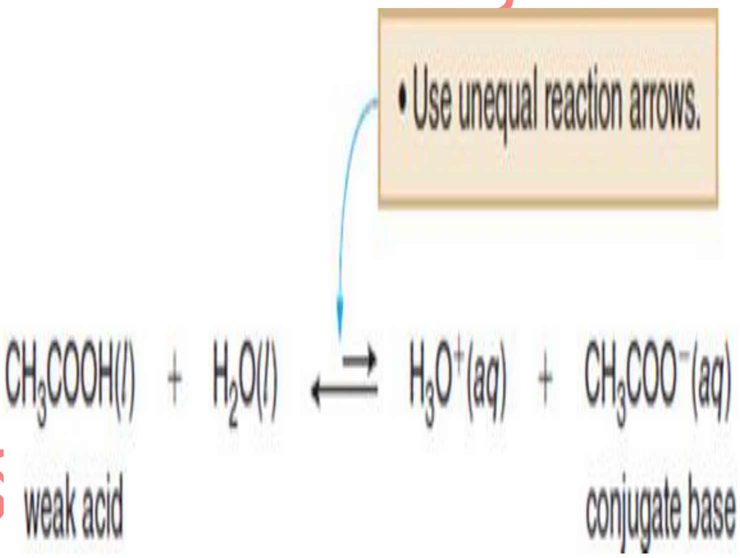
A weak acid less readily donates a proton.

When a weak acid dissolves in water, only a small

$HCl(g)$	+ $H_2O(l)$	$\longrightarrow$ $H_3O^+(aq)$	+	$Cl^-(aq)$
strong acid				conjugate base
$H_2SO_4(l)$	+ $H_2O(l)$	$\longrightarrow$ $H_3O^+(aq)$	+	$HSO_4^-(aq)$
strong acid				conjugate base



• Use unequal reaction arrows.

$$CH_3COOH(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$$

weak acid

conjugate base

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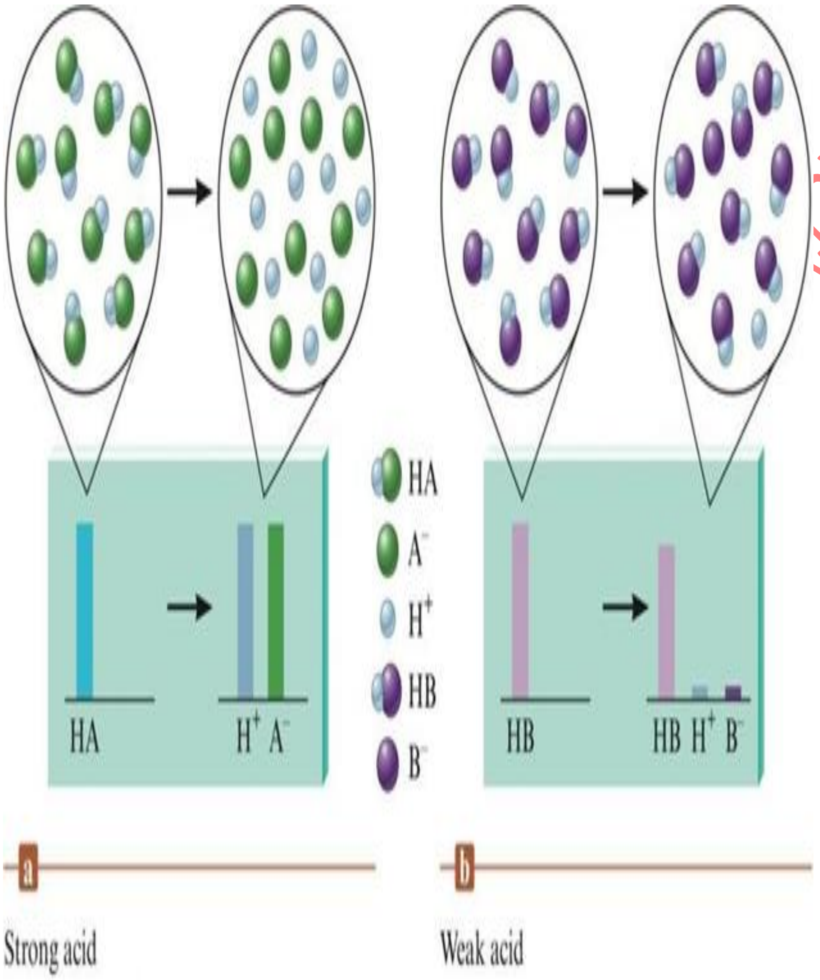

Fig.6.1: dissociation of the strong and weak acid

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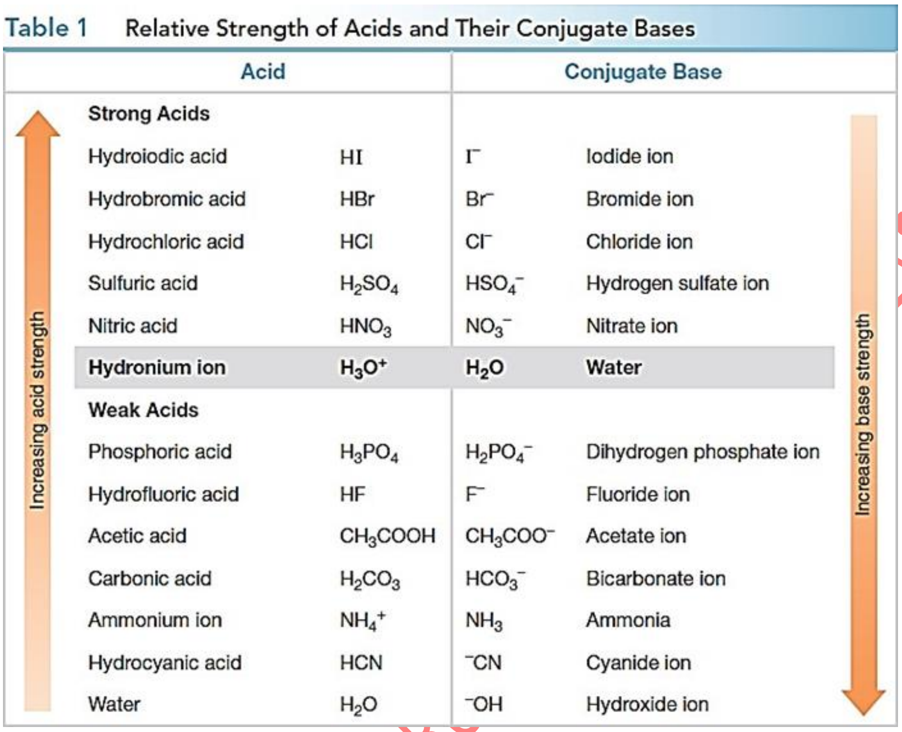

Table 1 Relative Strength of Acids and Their Conjugate Bases

Acid		Conjugate Base	
Strong Acids			
Hydroiodic acid	HI	I^-	Iodide ion
Hydrobromic acid	HBr	Br^-	Bromide ion
Hydrochloric acid	HCl	Cl^-	Chloride ion
Sulfuric acid	H_2SO_4	HSO_4^-	Hydrogen sulfate ion
Nitric acid	HNO_3	NO_3^-	Nitrate ion
Hydronium ion	H_3O^+	H_2O	**Water**
Weak Acids			
Phosphoric acid	H_3PO_4	$H_2PO_4^-$	Dihydrogen phosphate ion
Hydrofluoric acid	HF	F^-	Fluoride ion
Acetic acid	CH_3COOH	CH_3COO^-	Acetate ion
Carbonic acid	H_2CO_3	HCO_3^-	Bicarbonate ion
Ammonium ion	NH_4^+	NH_3	Ammonia
Hydrocyanic acid	HCN	^-CN	Cyanide ion
Water	H_2O	^-OH	Hydroxide ion

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Bases also differ in their ability to accept a proton.

A strong base readily accepts a proton.

When a strong base dissolves in water, 100% of the base dissociates into ions.

A weak base less readily accepts a proton. When a weak base dissolves in water, only a small fraction of the base forms ions.

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$$\mathrm{NaOH}(s) + \mathrm{H_2O}(l) \longrightarrow \mathrm{Na^+}(aq) + {}^{-}\mathrm{OH}(aq)$$

strong base

$$\mathrm{NH_3}(g) + \mathrm{H_2O}(l) \overset{\rightharpoonup}{\longleftarrow} \mathrm{NH_4^+}(aq) + {}^{-}\mathrm{OH}(aq)$$

weak base

- Use unequal reaction arrows.

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An inverse relationship exists between acid and base strength.

A strong acid readily donates a proton, forming a weak conjugate base.

A strong base readily accepts a proton, forming a weak conjugate acid.

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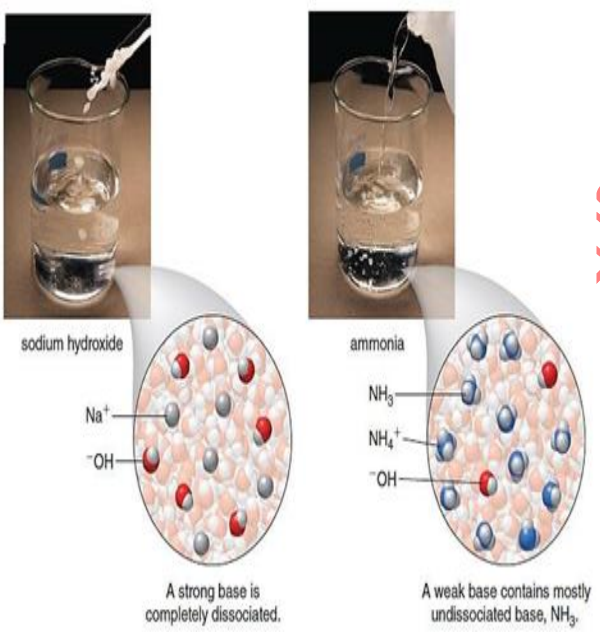

Fig.6.2: A Strong and Weak Base Dissolved in Water.

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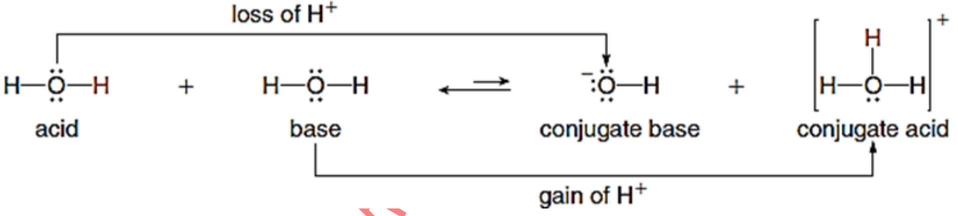
6-3Dissociation of Water

In Section5.1 we learned that water can behave as both a Bronsted–Lowry acid and a Bronsted–Lowry base.

As a result, two molecules of water can react together in an acid–base reaction.



One molecule of H_2O donates a proton .
(H^+), forming its conjugate base ^-OH

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**One molecule of H_2O .
accepts a proton, forming its
conjugate acid H_3O^+**

In pure water, $[H3O^+]$
$= [^-OH] = 1.0 \times 10^{-7}$ M

Multiplying these concentrations together gives the **ion–product constant** for water, symbolized by K_w.

$$\boxed{\underset{\text{ion–product constant}}{K_w} = [H_3O^+][^-OH]}$$

Substituting the concentrations for H_3O^+ and ^-OH into the expression for K_w gives the following result.

$$K_w = [H_3O^+][^-OH]$$
$$K_w = (1.0 \times 10^{-7}) \times (1.0 \times 10^{-7})$$