Conjugate acid

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A **conjugate acid**, within the Brønsted–Lowry acid–base theory, is a species formed by the reception of a proton (H^+) by a base—in other words, it is a base with a hydrogen ion added to it. On the other hand, a **conjugate base** is merely what is left after an acid has donated a proton in a chemical reaction. Hence, a conjugate base is a species formed by the removal of a proton from an acid.[1]

In summary, this can be represented as the following chemical reaction:

$$\text{Acid} + \text{Base} \rightleftharpoons \text{Conjugate Base} + \text{Conjugate Acid}$$

Johannes Nicolaus Brønsted and Martin Lowry introduced the Brønsted–Lowry theory, which proposed that any compound that can transfer a proton to any other compound is an acid, and the compound that accepts the proton is a base. A proton is a nuclear particle with a unit positive electrical charge; it is represented by the symbol H^+ because it constitutes the nucleus of a hydrogen atom,[2] that is, a hydrogen cation.

Johannes Nicolaus Brønsted (left) and Martin Lowry (right).

A cation can be a conjugate acid, and an anion can be a conjugate base, depending on which substance is involved and which acid–base theory is the viewpoint.

Contents

Acid-base reactions

In an acid-base reaction, an acid plus a base reacts to form a conjugate base plus a conjugate acid:

Conjugates are formed when an acid loses a hydrogen proton or a base gains a hydrogen proton. Refer to the following figure:

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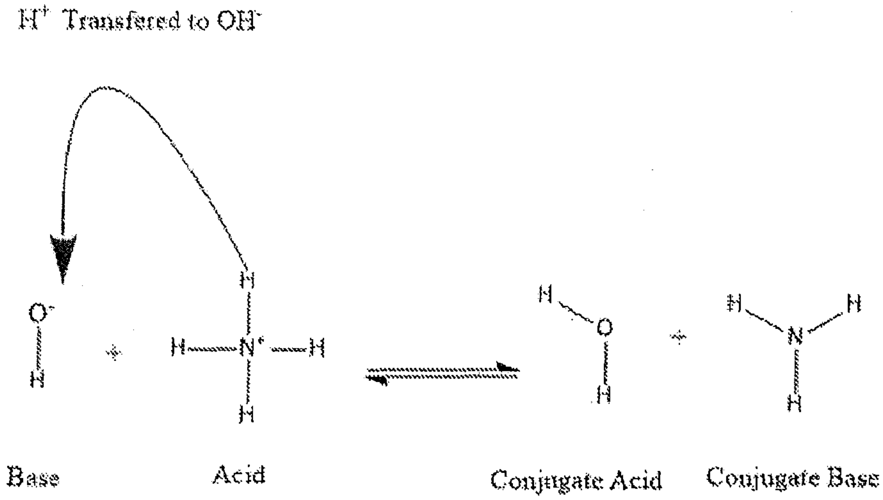

We say that the water molecule is the conjugate acid of the hydroxide ion after the latter received the hydrogen proton donated by ammonium. On the other hand, ammonia is the conjugate base for the acid ammonium after ammonium has donated a hydrogen ion towards the production of the water molecule. We can also refer to OH- as a conjugate base of H_2O, since the water molecule donates a proton towards the production of NH_4^+ in the reverse reaction, which is the predominating process in nature due to the strength of the base NH_3 over the hydroxide ion. Based on this information, it is clear that the terms "Acid", "Base", "conjugate acid", and "conjugate base" are not fixed for a certain chemical species; but are interchangeable according to the reaction taking place.

Strength of conjugates

The strength of a conjugate acid is directly proportional to its dissociation constant. If a conjugate acid is strong, its dissociation will have a higher equilibrium constant and the products of the reaction will be favored. The strength of a conjugate base can be seen as the tendency of the species to "pull" hydrogen protons towards itself. If a conjugate base is classified as strong, it will "hold on" to the hydrogen proton when in solution and its acid will not dissociate.

If a chemical species is classified as a weak acid, its conjugate base will be strong in nature. This can be observed in ammonia's (relatively strong base) reaction with water. The reaction proceeds until most of the ammonia has been transformed to ammonium. This shift to the right in the chemical equilibrium of the reaction means that ammonium does not dissociate easily in water (weak acid), and its conjugate base is stronger than the hydroxide ion.

On the other hand, if a species is classified as a strong acid, its conjugate base will be weak in nature. An example of this case would be the dissociation of Hydrochloric acid HCl in water. Since HCl is a strong acid (it dissociates to a great extent), its conjugate base (Cl^-) will be a weak conjugate base. Therefore, in this system, most H+ will be in the form of a Hydronium ion H_3O^+ instead of attached to a Cl anion and the conjugate base will be weaker than a water molecule.

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To summarize, the stronger the acid or base, the weaker the conjugate and vice versa.

Identifying conjugate acid-base pairs

The acid and conjugate base as well as the base and conjugate acid are known as conjugate pairs. When finding a conjugate acid or base, it is important to look at the reactants of the chemical equation. In this case, the reactants are the acids and bases, and the acid corresponds to the conjugate base on the product side of the chemical equation; as does the base to the conjugate acid on the product side of the equation.

To identify the conjugate acid, look for the pair of compounds that are related. The acid-base reaction can be viewed in a before and after sense. The before is the reactant side of the equation, the after is the product side of the equation. The conjugate acid in the after side of an equation gains a hydrogen ion, so in the before side of the equation the compound that has one less hydrogen ion of the conjugate acid is the base. The conjugate base in the after side of the equation lost a hydrogen ion, so in the before side of the equation the compound that has one more hydrogen ion of the conjugate base is the acid.

Consider the following acid-base reaction:

$$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$$

Nitric acid (HNO_3) is an *acid* because it donates a proton to the water molecule and its *conjugate base* is nitrate (NO_3^-). The water molecule acts as a base because it receives the Hydrogen Proton and its conjugate acid is the hydronium ion (H_3O^+).

Equation	Acid	Base	Conjugate Base	Conjugate Acid
$HClO_2 + H_2O \rightarrow ClO_2^- + H_3O^+$	$HClO_2$	H_2O	ClO_2^-	H_3O^+
$ClO^- + H_2O \rightarrow HClO + OH^-$	H_2O	ClO^-	OH^-	$HClO$
$HCl + H_2PO_4^- \rightarrow Cl^- + H_3PO_4$	HCl	$H_2PO_4^-$	Cl^-	H_3PO_4

Applications

One use of conjugate acids and bases lies in buffering systems, which include a buffer solution. In a buffer, a weak acid and its conjugate base (in the form of a salt), or a weak base and its conjugate acid are used in order to limit the pH change during a titration process. Buffers have both organic and non-organic chemical applications; for instance, besides buffers being used in lab processes, our blood acts as a buffer to maintain pH. The most important buffer in our bloodstream is the carbonic acid-bicarbonate buffer, which prevents drastic pH changes when CO_2 is introduced. This functions as such:

$$CO_2 + H_2O \rightleftarrows H_2CO_3 \rightleftarrows HCO_3^- + H^+$$

Furthermore, here is a table of common buffers.

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Buffering agent	pK_a	useful pH range
Citric acid	3.13, 4.76, 6.40	2.1 - 7.4
Acetic acid	4.8	3.8 - 5.8
KH_2PO_4,	7.2	6.2 - 8.2
CHES	9.3	8.3–10.3
Borate	9.24	8.25 - 10.25

A second common application with an organic compound would be the production of a buffer with acetic acid. If acetic acid, a weak acid with the formula CH_3COOH, was made into a buffer solution, it would need to be combined with its conjugate base CH_3COO^- in the form of a salt. The resulting mixture is called an acetate buffer, consisting of aqueous CH_3COOH and aqueous CH_3COONa. Acetic acid, along with many other weak acids, serve as useful components of buffers in different lab settings, each useful within their own pH range.

An example with an inorganic compound would be the medicinal use of lactic acid's conjugate base known as lactate in Lactated Ringer's solution and Hartmann's solution. Lactic acid has the formula $C_3H_6O_6$ and its conjugate base is used in intravenous fluids that consist of sodium and potassium cations along with lactate and chloride anions in solution with distilled water. These fluids are commonly isotonic in relation to human blood and are commonly used for spiking up the fluid level in a system after severe blood loss due to trauma, surgery, or burn injury.

Table of acids and their conjugate bases

Tabulated below are several examples of acids and their conjugate bases; notice how they differ by just one proton (H^+ ion). Acid strength decreases and conjugate base strength increases down the table.

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Acid	Conjugate Base
H_2F^+ Fluoronium ion	HF Hydrogen fluoride
HCl Hydrochloric acid	Cl^- Chloride ion
H_2SO_4 Sulfuric acid	HSO_4^- Hydrogen sulfate ion
HNO_3 Nitric acid	NO_3^- Nitrate ion
H_3O^+ Hydronium ion	H_2O Water
HSO_4^- Hydrogen sulfate ion	SO_4^{2-} Sulfate ion
H_3PO_4 Phosphoric acid	$H_2PO_4^-$ Dihydrogen phosphate ion
CH_3COOH Acetic acid	CH_3COO^- Acetate ion
H_2CO_3 Carbonic acid	HCO_3^- Hydrogen carbonate ion
H_2S Hydrosulfuric acid	HS^- Hydrogen sulfide ion
$H_2PO_4^-$ Dihydrogen phosphate ion	HPO_4^{2-} Hydrogen phosphate ion
NH_4^+ Ammonium ion	NH_3 Ammonia
HCO_3^- Hydrogencarbonate *(bicarbonate)* ion	CO_3^{2-} Carbonate ion
HPO_4^{2-} Hydrogen phosphate ion	PO_4^{3-} Phosphate ion
H_2O Water (neutral, pH 7)	OH^- Hydroxide ion

Table of bases and their conjugate acids

In contrast, here is a table of bases and their conjugate acids. Similarly, base strength decreases and conjugate acid strength increases down the table.

Base	Conjugate Acid
$C_2H_5NH_2$ Ethylamine	$C_2H_5NH_3^+$ Ethylammonium ion
CH_3NH_2 Methylamine	$CH_3NH_3^+$ Methylammonium ion
NH_3 Ammonia	NH_4^+ Ammonium ion
C_5H_5N Pyridine	$C_5H_6N^+$ Pyridinium
$C_6H_5NH_2$ Aniline	$C_6H_5NH_3^+$ Phenylammonium ion
$C_6H_5CO_2^-$ Benzoate ion	$C_6H_6CO_2$ Benzoic acid
F^- Fluoride ion	HF Hydrogen fluoride

See also

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