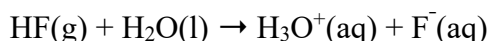


Acid and Bases

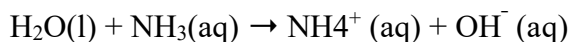
1. Bronsted Concept

Johannes Bronsted in Denmark and Thomas Lowry in England proposed (in 1923) that the essential feature of an acid–base reaction is the transfer of a hydrogen ion, H^+ , from one species to another. In the context of this definition, a hydrogen ion is often referred to as a proton. They suggested that any substance that acts as a proton donor should be classified as an acid, and any substance that acts as a proton acceptor should be classified as a base. Substances that act in this way are now called ‘Bronsted acids’ and ‘Bronsted bases’, respectively.

An example of a Bronsted acid is hydrogen fluoride, HF, which can donate a proton to another molecule, such as H_2O , when it dissolves in water:



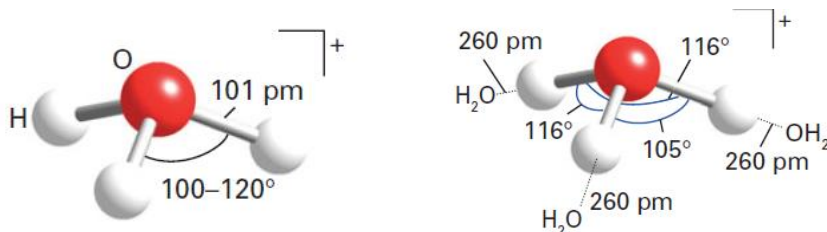
An example of a Bronsted base is ammonia, NH_3 , which can accept a proton from a proton donor:



As these two examples show, water is an example of an **amphiprotic** substance, a substance that can act as both a Bronsted acid and a Bronsted base.

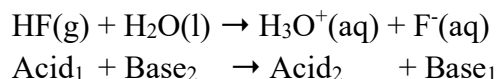
2. Structure of H^+ ion

When an acid donates a proton to a water molecule, the latter is converted into a *hydronium ion*, H_3O^+ (as per the crystal structure of $H_3O^+ClO_4^-$). However, the entity H_3O^+ is almost certainly an oversimplified description of the proton in water, for it participates in extensive hydrogen bonding, and a better representation is $H_9O_4^+(2)$.



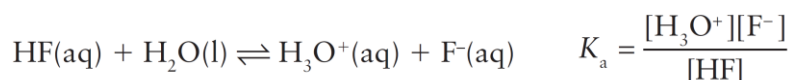
3. Conjugate acids and bases

The form of the two forward and reverse reactions given above, both of which depend on the transfer of a proton from an acid to a base, is expressed by writing the general Bronsted equilibrium as

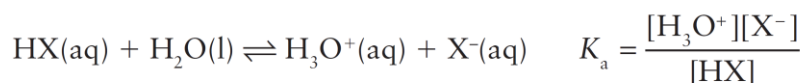


The species Base_1 is called the **conjugate base** of Acid_1 , and Acid_2 is the **conjugate acid** of Base_2 . The conjugate base of an acid is the species that is left after a proton is lost. The conjugate acid of a base is the species formed when a proton is gained. Thus, F^- is the conjugate base of HF and H_3O^+ is the conjugate acid of H_2O . There is no *fundamental* distinction between an acid and a conjugate acid or a base and a conjugate base: a conjugate acid is just another acid and a conjugate base is just another base.

4. The strengths of Bronsted acids



More generally:



In this definition, $[\text{X}^-]$ denotes the numerical value of the molar concentration of the species X^- (so, if the molar concentration of HF molecules is $0.001 \text{ mol dm}^{-3}$, then $[\text{HF}] = 0.001$).

A value $K_a \ll 1$ implies that $[\text{HX}]$ is large with respect to $[\text{X}^-]$, and so proton retention by the acid is favoured. The experimental value of K_a for hydrogen fluoride in water is 3.5×10^{-4} , indicating that under normal conditions only a very small fraction of HF molecules are deprotonated in water. The actual fraction deprotonated can be calculated as a function of acid concentration from the numerical value of K_a .

Numerical:

The pH of $0.145 \text{ M CH}_3\text{COOH(aq)}$ is 2.80. Calculate K_a of ethanoic acid.

Answer: To calculate K_a we need to calculate the concentrations of H_3O^+ , CH_3CO_2^- , and CH_3COOH in the solution. The concentration of H_3O^+ is obtained from the pH by writing $[\text{H}_3\text{O}^+] = 10^{\text{pH}}$, so in a solution of $\text{pH} = 2.80$, the molar concentration of H_3O^+ is $1.6 \times 10^{-3} \text{ mol dm}^{-3}$. Each deprotonation event produces one H_3O^+ ion and one CH_3CO_2^- ion, so the concentration of CH_3CO_2^- is the same as that of the H_3O^+ ions (provided the autoprotolysis of water can be neglected). The molar concentration of the remaining acid is $0.145 - 0.0016 \text{ mol dm}^{-3} = 0.143 \text{ mol dm}^{-3}$.

Therefore

$$K_a = \frac{(1.6 \times 10^{-3})^2}{0.143} = 1.8 \times 10^{-5}$$

This value corresponds to $pK_a = 4.75$.

Self-test For hydrofluoric acid $K_a = 3.5 \times 10^{-4}$. Calculate the pH of 0.10 M HF(aq).