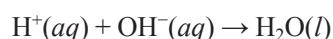
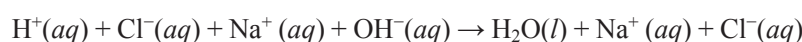
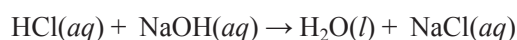


According to theory of the Swedish chemist **Arrhenius**, **acids** are those substances that ionise in water forming  $\text{H}^+$  ions, and **bases** are substances that in water produce  $\text{OH}^-$  ions.

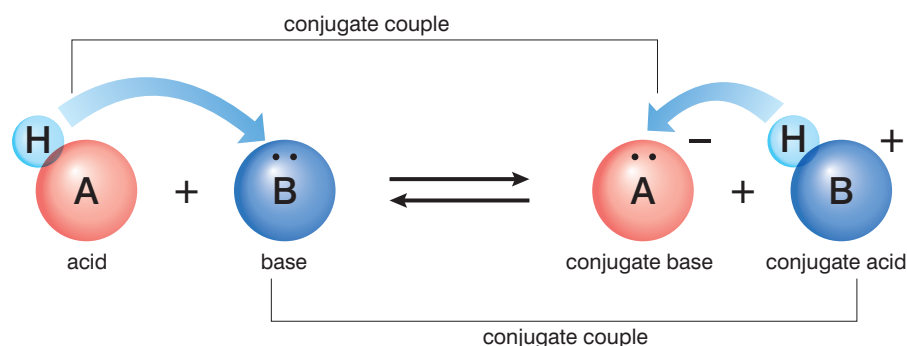
**Acid-base indicators** are substances that have the ability to take on a different colour depending on the pH of the solution in which they are placed.

Mixing an acid solution with a basic solution gives rise to a reaction called a **neutralisation reaction**, for which the simplified representation of the equation in ionic form is the same regardless of the acid or base:  $\text{H}^+$  ions from the acid and  $\text{OH}^-$  ions from the base combine to form  $\text{H}_2\text{O}$  molecules; anions from the acid and cations from the base produce the other reaction product, a salt, which varies depending on the type of acid or base.

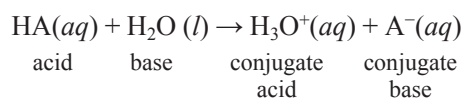
For example:



According to the **theory of Brønsted and Lowry** the concept of acid and base is no longer limited to aqueous solutions, the definition becomes *acids are all chemical species that donate protons and bases are chemical species that acquire protons*.



An important observation that stems directly from these definitions is the following: a chemical species behaves like an acid only when in the presence of another chemical species that acts as a base. It follows that all acid-base reactions consist of a proton exchange between the acid and base. Also, the ionisation reaction in water of a generic acid HA is an acid-base reaction and can be represented by the following equation:



From this equation the expression for the **acid ionisation constant (K<sub>a</sub>)** is derived:

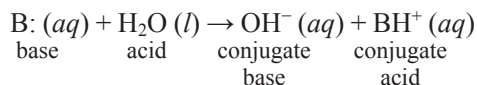
$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

The value of the ionisation equilibrium constant  $K_a$  represents the strength of an acid: the stronger an acid is the greater is its tendency to donate a proton to the water and the higher the value of  $K_a$ .

The values of  $K_a$  help interpret how the equilibrium of a reaction has shifted: all acid-base equilibria are shifted in the direction in which the weaker acid and base form.

*Monoprotic acids* are defined as those molecules which are able to donate just one proton per molecule; acids that can donate more than one proton per molecule in general are called *polyprotic acids*.

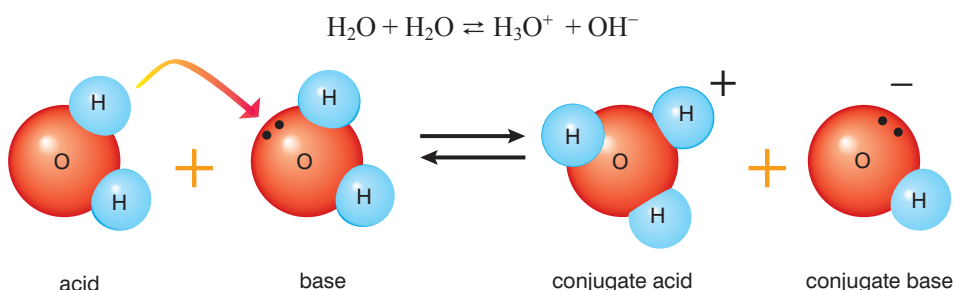
Similarly, a base is stronger the greater is its tendency to acquire a proton from water, and the higher is the value of its **basic dissociation constant (K<sub>b</sub>)**.



The expression for the constant **K<sub>b</sub>** is:

$$K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{:B}]}$$

In addition to determining the ionisation of many substances, water has the unique feature of ionising itself. The equilibrium for the **self-ionisation of water** is represented by the equation:



Equilibrium is strongly shifted to the left in order to keep the concentration of water molecules constant. At a temperature of 25 °C, the product of the concentration of  $\text{H}_3\text{O}^+$  ion and that of  $\text{OH}^-$  ions is constant and is always  $1 \cdot 10^{-14}$ .

The **ionic product of water** constant is denoted by **K<sub>w</sub>**:

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \cdot 10^{-14}$$

To describe in simple terms the case of neutrality and all other situations in which the system is no longer neutral the **pH** scale is used, which originates from the definition:

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

Therefore, at 25 °C a solution is acidic if the pH is less than 7: the higher the pH value is the less acidic the solution is.

A solution is basic if the pH is greater than 7: the higher the pH value the more basic the solution is.

At  $\text{pH} = 7$  the system is neutral.

To determine the pH value without making calculations it is useful to remember that if the ion concentration of  $\text{H}_3\text{O}^+$  is expressed simply as a power of ten, the pH value is equal to the exponent of the power but of opposite sign.

If two solutions are compared, with the same molar concentration but prepared with different acid strength, it is easily seen that the pH of the solution of the weaker acid is higher than that of the strong acid: in the solution of the weaker acid the concentration of  $\text{H}_3\text{O}^+$  ions is lower.

Whenever solutions of a strong acid and a strong base are mixed in stoichiometrically equivalent amounts a neutral solution ( $\text{pH} = 7$ ) results.

As with strong acids, weak acids also react with a stoichiometric amount of base; in this case, however, the solution obtained has a pH greater than 7 because the solutions of salts consisting of an anion derived from a weak acid and a cation that comes from a strong base exhibit a behaviour that is referred to as **basic hydrolysis**.

Similarly, when a weak base is neutralised by a strong acid a solution of acidic character is obtained, in fact, because of **acid hydrolysis**, solutions of salts are acidic as they are composed of an anion derived from a strong acid and a cation derived from a weak base.

Solutions consisting of a *mixture of a weak base with its conjugate acid* (for example  $\text{NH}_3$  and  $\text{NH}_4^+$ ) or a *mixture of a weak acid and its conjugate base* (for example  $\text{CH}_3\text{COOH}$  and  $\text{CH}_3\text{COO}^-$ ) have the property of maintaining a virtually unchanged pH value, even if small amounts of strong acids or strong bases are added: these systems are called **buffer solutions**.