Acid dissociation constant or acidity constant (symbol: $K_a$) of the hypothetical compound HA is the equilibrium constant of the reaction (see, Brønsted-Lowry theory),

$$\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{A}^- + \text{H}_3\text{O}^+$$

in dilute aqueous solution at 25 ºC.

For the above system,

$$\frac{[\text{A}^-] [\text{H}_3\text{O}^+]}{[\text{HA}] [\text{H}_2\text{O}]} = \text{constant (K)}$$

In dilute solution,

$$[\text{H}_2\text{O}] \sim \text{large (} \sim 56 \text{ molL}^{-1} \sim \text{constant (k)})$$

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

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

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

$$K_a (\text{acid}) = \frac{[\text{conjugate base}] [\text{H}_3\text{O}^+]}{[\text{acid}]}$$

The larger the $K_a$ of a species, the larger the fraction of the species that donates H+ to water, greater its acid strength. Thus, $K_a$ of a species is a measure of its acid strength.

eg:

<table>
<thead>
<tr>
<th>compound</th>
<th>$K_a$</th>
</tr>
</thead>
<tbody>
<tr>
<td>HI</td>
<td>$10^{10}$</td>
</tr>
<tr>
<td>HCl</td>
<td>$10^7$</td>
</tr>
</tbody>
</table>

$K_a$ of HI is larger than that of HCl, meaning HI is a stronger acid than HCl.

Contributors

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