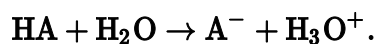


PH of strong acids and bases

When calculating pH, it is always necessary to consider what is the source of oxonium cations in a given environment.

Strong monosaturated acids

For **strong monosaturated acids** the dissociation follows the equation



For the calculation we assume:

- the substance quantity of H_3O^+ according to the above equation will be the same as A^- , which, given an identical volume, is also true for the concentration, i.e. $[\text{H}_3\text{O}^+] = [\text{A}^-]$;
- all acid - because it is a strong acid - is converted into A^- a H_3O^+ , therefore we will mark $[\text{A}^-]$ its concentration, i.e. $[\text{A}^-] = c_{\text{HA}}$

Let's deduce:

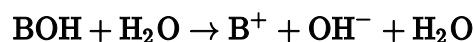
$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log [\text{A}^-] = -\log c_{\text{HA}},$$

and to calculate the pH we get the formula

$$\text{pH} = -\log c_{\text{HA}}.$$

Strong monosaturated bases

For **strong monosaturated bases** the dissociation follows the equation



We assume, as in the case of strong monosaturates, that:

- the amount, or concentration, of hydroxide ions and the resulting B^+ is the same according to the above chemical equation, i.e. $[\text{OH}^-] = [\text{B}^+]$;
- dissociation occurs completely, i.e. $[\text{B}^+] = c_{\text{BOH}}$.

The calculation is therefore analogous, we just have to remember that unlike acids, the base is not a source of oxonium cations, but takes oxonium cations from the environment (see the theory of acids and bases), so we add from the equation for the ionic product of water:

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]}$$

and from these assumptions, we deduce

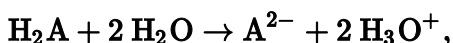
$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log \frac{K_w}{[\text{OH}^-]} = \log [\text{OH}^-] - \log K_w = \log [\text{B}^+] - \log K_w = \log c_{\text{BOH}} - \log K_w.$$

Calculate the pH at 25 °C using the formula

$$\text{pH} = 14 + \log c_{\text{BOH}}.$$

Strong dibasic acids

Strong dibasic acids dissociate according to the equation



we assume, then:

- complete dissociation, i.e. $c_{\text{H}_2\text{A}} = [\text{A}^{2-}]$;
- however, the amount of oxonium cations and the amount of formed A^{2-} is - in contrast to monosaturated acids - in a ratio of 1:2, i.e. $[\text{H}_3\text{O}^+] = 2 \cdot c_{\text{H}_2\text{A}}$.

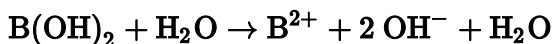
From this we derive $\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log [\text{A}^{2-}] = -\log(2 \cdot c_{\text{H}_2\text{A}}) = -\log 2 - \log c_{\text{H}_2\text{A}}$

and the pH is calculated according to the formula

$$\text{pH} = -\log c_{\text{H}_2\text{A}} - \log 2.$$

Strong dibasic bases

Strong dibasic bases dissociate according to the equation



as we assume for monosaturated bases and dibasic acids:

- complete dissociation, i.e. $c_{\text{B(OH)}_2} = [\text{B}^{2+}]$;
- concentration of the formed B^{2+} and the concentration of hydroxide anions is in the ratio 1:2, i.e. $[\text{OH}^-] = 2 \cdot [\text{B}^{2+}]$, in addition, according to the previous assumption $[\text{OH}^-] = 2 \cdot c_{\text{B(OH)}_2}$
- hydroxide anions drain oxonium cations from the environment, $[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]}$.

Then we derive

$$-\log[\text{H}_3\text{O}^+] = -\log \frac{K_w}{[\text{OH}^-]} = \log [\text{OH}^-] - \log K_w = \log(2 \cdot c_{\text{B(OH)}_2}) - \log K_w = \log 2 + \log c_{\text{B(OH)}_2} - \log K_w$$

and the pH at 25 °C is calculated according to the formula

$$\text{pH} = 14 + \log 2 + \log c_{\text{B(OH)}_2}.$$

References

Related articles

- pH of weak acids and bases
- pH-metry
- pH measurment
- Buffers, buffering capacity, oxidation and reduction, electrode processes
- pH of urine
- pH of salts