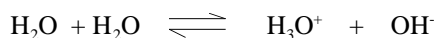


ACID/BASE THEORY

For the ionization of water :



an equilibrium expression can be written:

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14} \quad (@ 25^\circ\text{C})$$

— or —

$$K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14} \quad (@ 25^\circ\text{C})$$

K_w is known as the **ion product** of water. In **ANY** aqueous solution this equilibrium is always present. One can always obtain the concentration of H^+ from the concentration of OH^- and vice versa. A strong acid will set the value of $[\text{H}^+]$ and then $[\text{OH}^-]$ can be calculated. A strong base will set the value of $[\text{OH}^-]$ and then $[\text{H}^+]$ can be calculated.

The pH of a solution can be determined by the expression: $\text{pH} = -\log[\text{H}^+]$ which means that $[\text{H}^+] = 10^{-\text{pH}}$

In general, the "p" in pH is really a function: $\text{p}(X) = -\log(X)$ but instead of writing $\text{p}(X)$ we write $\text{p}X$.

So you can easily calculate values for pOH, $\text{p}K_w$, etc...

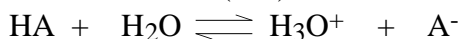
This leads to the "-log" form of the above equation: $14 = \text{pH} + \text{pOH}$

The **ONLY** time that $[\text{H}^+] = [\text{OH}^-]$ (definition of neutral water) is when they each equal $1.0 \times 10^{-7} \text{ M}$ which corresponds to a pH (and pOH for that matter) of 7.00. This is neutral pH and is the pH of pure H_2O (note this is only at 25°C , however). pH values lower than 7 are acidic solutions while pH values higher than 7 are basic.

NOTE : ALL the questions in chapters 6 and 7 will fit into these equilibria and equations compare and contrast the similarities and differences in these two columns

— ACIDS —

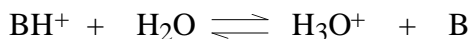
For a weak acid (HA) in water:



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

$$\text{pH} = \text{p}K_a + \log \frac{[\text{A}^-]}{[\text{HA}]}$$

For a conjugate acid (BH^+) of a weak base (B) in water:



$$K_a = \frac{[\text{H}^+][\text{B}]}{[\text{BH}^+]} \quad [\text{H}^+] = K_a \frac{[\text{BH}^+]}{[\text{B}]}$$

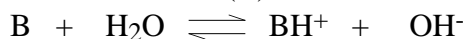
$$\text{pH} = \text{p}K_a + \log \frac{[\text{B}]}{[\text{BH}^+]}$$

NOTICE! You can NOT look up the K_a of BH^+ . You have to calculate it from the K_b of B:

$$K_a = \frac{K_w}{K_b}$$

— BASES —

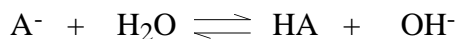
For a weak base (B) in water:



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

$$\text{pOH} = \text{p}K_b + \log \frac{[\text{BH}^+]}{[\text{B}]}$$

For a conjugate base (A^-) of a weak acid (HA) in water:



$$K_b = \frac{[\text{OH}^-][\text{HA}]}{[\text{A}^-]} \quad [\text{OH}^-] = K_b \frac{[\text{A}^-]}{[\text{HA}]}$$

$$\text{pOH} = \text{p}K_b + \log \frac{[\text{HA}]}{[\text{A}^-]}$$

NOTICE! You can NOT look up the K_b of A^- . You have to calculate it from the K_a of HA:

$$K_b = \frac{K_w}{K_a}$$