



The Basics of Acids and Bases

Acids:

The two main definitions of acids are:

- Bronsted-Lowry Acid: Molecules that can donate a hydrogen ion (proton donors)
- Lewis Acid: Molecules that can accept an electron pair (electron acceptors)

Bases:

- Bronsted-Lowry Base: Molecules that can accept hydrogen ions (proton acceptors)
- Lewis Base: Molecules that can donate electron pairs (electron donors)

pH, pOH, pKw :

- pH is a measure of the concentration of hydronium ions

$$pH = -\log([H_3O^+])$$

$$[H_3O^+] = 10^{-pH}$$

- pOH is a measure of the concentration of hydroxide ions

$$pOH = -\log([OH^-])$$

$$[OH^-] = 10^{-pOH}$$

- The p in front of the H and OH indicates a $-\log$. This stands to reason that pKw is:

$$pKw = -\log(Kw)$$

- The relationship between Kw , hydronium concentration, and hydroxide concentration is:

$$Kw = 1 \times 10^{-14} = [H_3O^+][OH^-]$$

- If we take the $-\log$ of everything, we can develop the full relationship

$$-\log(K_w) = -\log(1 \times 10^{-14}) = -\log([H_3O^+][OH^-])$$

$$pK_w = 14 = -\log([H_3O^+]) + (-\log([OH^-])) = pH + pOH$$

Examples

- Calculate the $[H_3O^+]$, $[OH^-]$, pH , and pOH for .1M HNO_3
 - Since HNO_3 is a strong acid, all the hydrogen ions will be dissociated in water to form H_3O^+

$$[H_3O^+] = .1M$$

$$pH = -\log([H_3O^+]) = -\log(.1) = 1$$

$$pOH = pK_a - pH = 14 - 1 = 13$$

$$[OH^-] = 10^{-pOH} = 10^{-13}$$

- Calculate the $[H_3O^+]$, $[OH^-]$, pH , and pOH for .05M $NaOH$
 - This works the same way, but $NaOH$ is a strong base, so we can find the concentration of OH^- first and work backwards to find the concentration of H_3O^+

$$[OH^-] = [NaOH] = .05M$$

$$pOH = -\log([OH^-]) = -\log(.05) = 1.30$$

$$pH = 14 - pOH = 14 - 1.3 = 12.70$$

$$[H_3O^+] = 10^{-12.7} = 2 \times 10^{-13}$$

For more information on acids and bases, make an appointment with a [content tutor](#).

Appointments are available in person at the [Student Success Center](#) or online.