

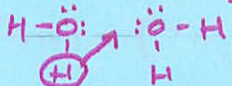
0.0000001 M

← same →

Notes 20.2

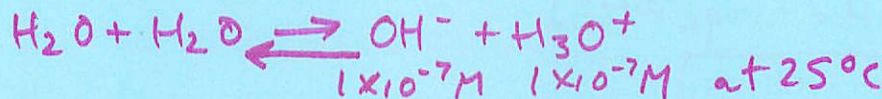
pH = pouvoir hydrogen = hydrogen power $[H^+]$ or $[H_3O^+]$

Self-Ionization of Water



at 25°C

Equation:



The product of the concentration of hydronium and hydroxide ions is constant in water and dilute acid/base solutions.

Formula: $K_w = [H_3O^+][OH^-]$ $[] = \text{molarity}$ $K_w = 1 \times 10^{-14} M^2$

↑ ionization constant of H_2O

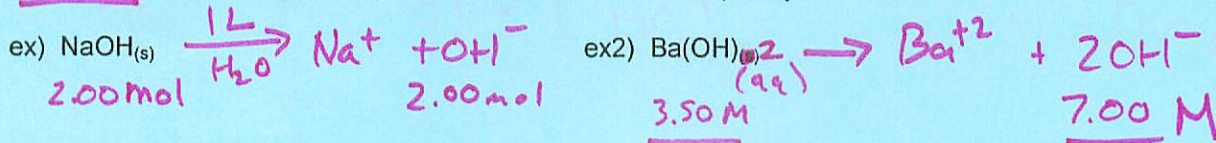
at 25 C: pure H_2O $[1 \times 10^{-7} M][1 \times 10^{-7} M] =$

Acid Soln: $[H_3O^+] > [OH^-]$

Base Soln: $[H_3O^+] < [OH^-]$

Neutral Soln: $[H_3O^+] = [OH^-]$

Calculations - remember, strong acids and bases ionize completely:



If we dissolve 1.0 mol of NaOH in 10.0 L of soln:

$$M = \frac{1.0 \text{ mol}}{10.0 L} = \boxed{0.10 M NaOH} = 0.10 M OH^-$$

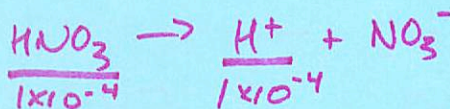
Using $K_w = [H_3O^+][OH^-]$ to find $[H_3O^+]$

$$\frac{K_w}{[OH^-]} = [H_3O^+] \quad \frac{1 \times 10^{-14} M^2}{0.10 M (1 \times 10^{-1})} = 1 \times 10^{-13} M [H_3O^+]$$

-remember, you subtract

-notice inverse relationship

You try: Determine hydronium and hydroxide concentrations of a $1.0 \times 10^{-4} M HNO_3$ solution:



$$K_w = [H_3O^+][OH^-]$$

$$\frac{K_w}{[H_3O^+]} = [OH^-] \quad \frac{1 \times 10^{-14}}{1 \times 10^{-4}} = 1 \times 10^{-10} M OH^-$$

An aqueous solution of barium hydroxide has a hydronium concentration of $1.0 \times 10^{-11} M$. What is the hydroxide molarity? What is the molarity of the $Ba(OH)_2$ solution?

$$K_w = [H_3O^+][OH^-]$$

$$1 \times 10^{-14} = [1 \times 10^{-11}][OH^-]$$

$$1 \times 10^{-3} \frac{M}{L} OH^- \left(\frac{1 \text{ mol } Ba(OH)_2}{2 \text{ mol } OH^-} \right) = 5 \times 10^{-4} M Ba(OH)_2$$

pH scale

pH formula:

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

log = power to which 10 must be raised to equal the number

$$\text{neutral pH} = [H_3O^+] = 1 \times 10^{-7} = 7$$

pOH formula:

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

Handy relationship:

$$pH + pOH = 14$$

ex) Find pH of a 1.0×10^{-3} M NaOH solution:

$$-\log[OH^-] = -\log[1 \times 10^{-3}] = 3 = pOH$$

$$3 + \frac{\quad}{pH} = 14$$

$$pH = 11 \quad \text{Bases } pH > 7$$

ex2) Find pH of a 3.4×10^{-5} M HCl solution

$$[H_3O^+] = 3.4 \times 10^{-5} \quad -\log[3.4 \times 10^{-5}] = 4.5 \text{ Acid } < 7$$

Calculating $[H_3O^+]$ and $[OH^-]$ from pH

math:

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

$$[OH^-] = 10^{(-pOH)} \quad 10^x$$

ex) Find $[H_3O^+]$ of a solution with a pH of 4.0.

$$1 \times 10^{-4} = [H_3O^+]$$

ex2) The pH of a solution is measured and determined to be 7.52.

Find pOH

$$7.52 + \frac{\quad}{\quad} = 14 \quad pOH = 6.48$$

Find $[H_3O^+]$

$$[H_3O^+] = 10^{(-7.52)} = 3.0 \times 10^{-8} M$$

Find $[OH^-]$

$$[H^+][OH^-] = \frac{1 \times 10^{-14}}{3.0 \times 10^{-8}} = 3.3 \times 10^{-7} M OH^-$$

Acid or Base? Base