0.0000001 M

same

Notes 20.2 $\boxed{pH}$ = pouvoir hydrogen = hydrogen power $[H^+]$ or $[H_3O^+]$

Self-Ionization of Water

H–Ö: :Ö–H
(+1) H

at 25°c

Equation:

$$H_2O + H_2O \rightleftharpoons OH^- + H_3O^+$$

$1\times10^{-7}M$ $1\times10^{-7}M$ at 25°c

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

Formula: $K_w = [H_3O^+][OH^-]$ $[\ ]$ = molarity $\boxed{K_w = 1\times10^{-14}M^2}$

↑ ionization constant of H_2O

at 25 C: pure H_2O $[1\times10^{-7}M][1\times10^{-7}M]$ =

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

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

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

Calculations - remember, strong acids and bases ionize completely:

ex) $NaOH_{(s)} \xrightarrow[H_2O]{1L} Na^+ + OH^-$
2.00 mol 2.00 mol

ex2) $Ba(OH)_{2(aq)} \longrightarrow Ba^{+2} + 2OH^-$
3.50 M 7.00 M

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

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

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

$$\frac{K_w}{[OH^-]} = [H_3O^+] \quad \frac{1\times10^{-14}M^2}{0.10\ M\ (1\times10^{-1})} = 1\times10^{-13}\ M\ [H_3O^+]$$

-remember, you subtract

-notice inverse relationship

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

$HNO_3 \rightarrow H^+ + NO_3^-$
1×10^{-4} 1×10^{-4}

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

$$\frac{K_w}{[H_3O^+]} = OH^- \qquad \frac{1\times10^{-14}}{1\times10^{-4}} = 1\times10^{-10}\ M\ OH^-$$

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

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

$1\times10^{-14} = [1\times10^{-11}][1\times10^{-3}]$

$$1\times10^{-3}\ \frac{Mol\ OH^-}{L}\left(\frac{1\ mol\ Ba(OH)_2}{2\ mol\ OH^-}\right) = 5\times10^{-4}\ M\ Ba(OH)_2$$

pH scale

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

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

neutral pH = $[H_3O^+] = 1\times10^{-7} = 7$

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

Handy relationship: $\boxed{pH + pOH = 14}$

ex) Find pH of a 1.0×10^{-3} M NaOH solution: $-\log[OH^-] = -\log[1\times10^{-3}] = 3 = pOH$

$3 + \underset{pH}{\underline{\quad}} = 14$ $pH = 11$ Bases pH > 7

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

$[H_3O^+] = 3.4\times10^{-5}$ $-\log[3.4\times10^{-5}] = 4.5$ ACID < 7

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

math: $[H_3O^+] = 10^{(-pH)}$ $[OH^-] = 10^{(-pOH)}$ 10^x

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

$1\times10^{-4} = [H_3O^+]$

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

Find pOH $7.52 + \underline{\quad} = 14$ $pOH = 6.48$

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

Find $[OH^-]$ $[H^+][OH^-] = \dfrac{1\times10^{-14}}{3.0\times10^{-8}} = 3.3\times10^{-7}$ M OH^-

Acid or Base? Base