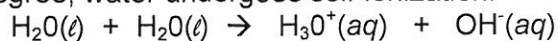


### pH Calculations

To a small but measurable degree, water undergoes self-ionization:



The use of brackets is a way of abbreviating "molar concentration." Thus,  $[\text{H}^+]$  can be read, "the concentration of hydrogen ion in moles per liter.  $[\text{OH}^-]$  can be read, "the concentration of hydroxide ion in moles per liter.

At 25°C,  $[\text{H}^+] = 1.0 \times 10^{-7}$  moles  $\text{H}^+$  per liter of solution  
 $[\text{OH}^-] = 1.0 \times 10^{-7}$  moles  $\text{OH}^-$  per liter of solution

Ionization Constant for Water ( $K_w$ )

1.  $K_w = [\text{H}^+][\text{OH}^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) = 1.0 \times 10^{-14}$
2.  $K_w$  is a constant at ordinary ranges of room temperatures

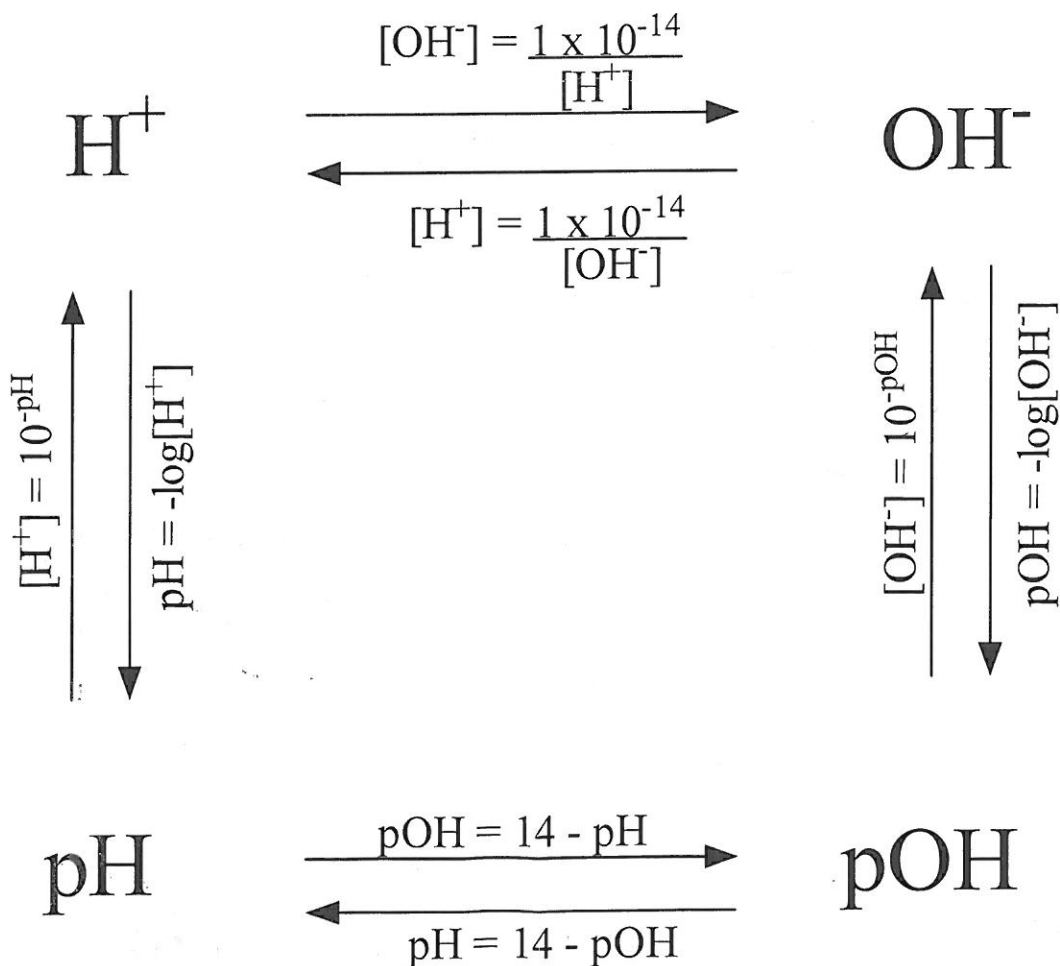
pH is the negative of the common logarithm of the hydrogen ion concentration

$$\text{pH} = -\log [\text{H}^+]$$

pOH is the negative of the common logarithm of the hydroxide ion concentration

$$\text{pOH} = -\log [\text{OH}^-]$$

$$\text{pH} + \text{pOH} = 14.0$$



$$[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

$$\text{pH} + \text{pOH} = 14$$

Acids  $\text{pH} = -\log [\text{H}_3\text{O}^+]$  Bases  $\text{pOH} = -\log [\text{OH}^-]$

$\text{pH} = 2$  then  $[\text{H}_3\text{O}^+] = 1.0 \times 10^{-2} \text{ M}$

$\text{pOH} = 12$  then  $[\text{OH}^-] = 1.0 \times 10^{-12} \text{ M}$

$\text{pH} = 3.4$  then  $[\text{H}_3\text{O}^+] = 10^{-3.4} = 3.98 \times 10^{-4} \text{ M}$

Now find

$[\text{OH}^-] = ?$  one way: Find pOH

$\text{pOH} = 14 - 3.4 = 10.6 \quad \therefore [\text{OH}^-] = 10^{-10.6} = 2.51 \times 10^{-11} \text{ M}$

Another way:

$$[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

$$\frac{3.98 \times 10^{-4} [\text{OH}^-]}{3.98 \times 10^{-4}} = \frac{1.0 \times 10^{-14}}{3.98 \times 10^{-4}} = 2.51 \times 10^{-11} \text{ M}$$