pH and pOH

- when working with dilute solutions of strong acids or weak acids (likewise with bases)
  > $[H_3O^+]$ or $[OH^-]$ is very small, often around 10^{-6} or smaller
  > with such small quantities, it is difficult to compare concentrations
  > pH and pOH scales developed

- pH = negative logarithm of the molar concentration of hydronium ion

- pOH = negative logarithm of molar concentration of hydroxide ion

$$pH = -\log[H3O^+]$$

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

$[H_3O^+] = 0.00025M$

$pH = 3.6$

logarithm = the exponent to which 10 must be raised to represent a certain number

For example . . . log100 **means** 10 to the power of WHAT equals 100?

$\log 100 = 2$ because $10^2 = 100$

$\log 1000 = 3$ $1000 = 10^3$

$\log(0.01) = -2$ 10^{-2}

The reverse procedure of log is called **antilog**.

Antilog is equivalent of raising 10 to the power of a certain exponent.

$$\text{antilog}(x) = 10^x$$

What is the antilog of 2?

$\text{antilog}(2) = 10^2 = 100$

Examples:

a) If the $[H_3O^+] = 4.67 \times 10^{-5}$ M, what is the pH?

$$pH = -\log[H_3O^+]$$
$$= -\log(4.67 \times 10^{-5})$$
$$= 4.331$$

Note: in pH and pOH, only the numbers after the decimal are significant

b) If the $[OH^-] = 2.83 \times 10^{-6}$ M, what is the pOH?

$$pOH = -\log[OH^-]$$
$$= 5.548$$

Consider the following:

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

- if we want to rearrange the equation to solve for $[H_3O^+]$, first remove the negative sign by multiplying both sides by -1 to give:

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

- since the reverse of log is antilog, the $[H_3O^+]$ is taking the antilog, the $[H_3O^+]$ can be determined by taking the antilog of both sides of the equation:

$$\text{antilog}(-pH) = \text{antilog}(\log[H_3O^+]) = [H_3O^+]$$

$$\text{antilog}(-pH) = [H_3O^+]$$

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

- similarly, $[OH^-]$ can be determined from the pOH

$$[H_3O^+] = \text{antilog}(-pH) = 10^{-pH}$$
$$[OH^-] = \text{antilog}(-pOH) = 10^{-pOH}$$

Example:

a) If the pH is 3.17, what is the $[H_3O^+]$?

$[H_3O^+] = 10^{-3.17}$ $\Rightarrow$ 6.8×10^{-4} M

$= \text{antilog} -3.17$

b) If the pOH is 5.32, what is the $[OH^-]$?

$[OH^-] = 10^{-5.32}$

$= \text{antilog} -5.32$ } 4.8×10^{-6} M

There is a simple but important relationship between pH, pOH and Kw.

- starting with Kw:

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

- take the logarithm of both sides:

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

$$\log[H_3O^+] + \log[OH^-] = \log Kw$$

- multiply by -1:

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

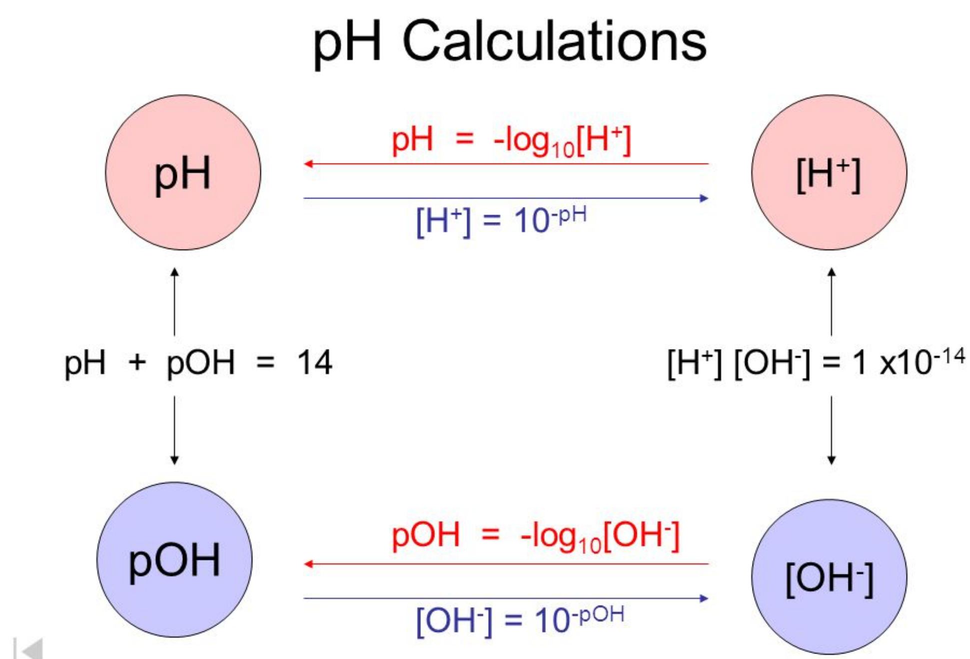
$$pH + pOH = pKw$$

- at 25 Kw = 1.0×10^{-14}

$$pH + pOH = -\log(1.0 \text{ X } 10^{-14}) = 14.00$$

$$\boxed{pH + pOH = 14.00}$$

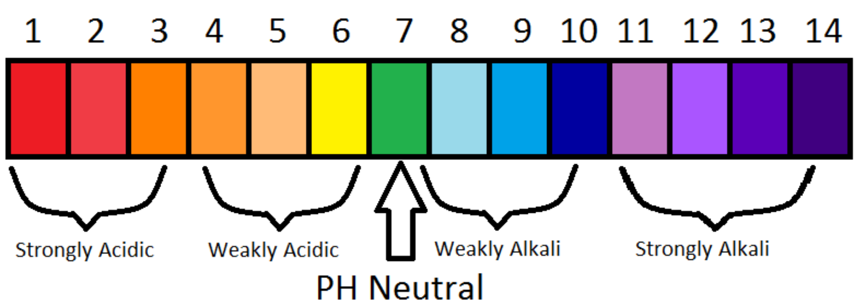
- Using the following relationships, you can work back and forth from any of $[H_3O^+]$, $[OH^-]$, pH and pOH.



- since the pH and pOH scales are logarithmic scales, a difference in one pH or pOH unit is equivalent to a ten-fold (10×) difference in concentration

  > pH and pOH are the negative of the exponent, so low pH and pOH mean relatively high values of $[H_3O^+]$ and $[OH^-]$ & high values of pH and pOH mean relatively low $[H_3O^+]$ and $[OH^-]$

at 25°C: pH = 7.0 is neutral
pH < 7.0 is acidic
pH > 7.0 is basic (alkaline)



- negative pH values are possible for concentrated strong acids

-log 2

2.00 M HCl has $[H_3O^+]$ = 2.00 mol/L

$$pH = -\log(2.00) = -0.30$$