ACIDITY AND ALKALINITY

- At pH = 7, pH = pOH. The solution is considered NEUTRAL

Also, $[H^+] = [OH^-]$!

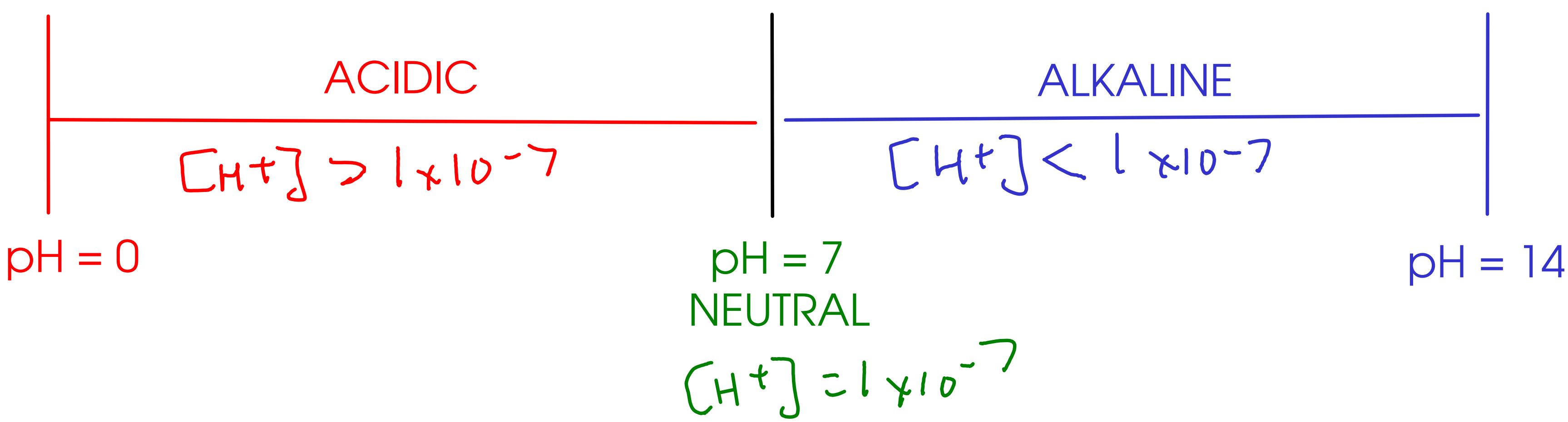
- At pH <7, pH < pOH. The solution is considered ACIDIC

Also, $[H^+] > [OH^-]$!

- At pH > 7, pH > pOH. The solution is considered ALKALINE (BASIC)

Also, $[H^+] < [OH^-]$!

The pH scale...



pH AND TEMPERATURE

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

This equation is valid at room temperature, specifically 25°C.

Equilibrium constants depend on TEMPERATURE, and change with temperature.

So, the "neutral" pH (where the concentration of hydroxide and hydronium ions are equal) CHANGES with changing temperatures

This change is important at temperatures greatly different from 25° C.

As an example, consider average "normal" human body temperature: 37° C

At 37°C, $pK_w = 13.60$

pH of neutral solution = $\underline{\underline{6.8}}$

ACID-BASE EQUILIBRIUM IN WATER

- Like other ELECTROLYTES, acids and bases IONIZE to some extent in water

- STRONG electrolytes ionize completely. Acids and bases that ionize completely in water are called STRONG ACIDS and STRONG BASES

- WEAK electrolytes ionize partially, remaining mostly non-ionized. Acids and bases that ionize only partially in solution are called WEAK ACIDS and WEAK BASES.

- Most acids and bases are WEAK!

Common strong acids

HCl

HNO_3

H_2SO_4 (only 1st proton)

HBr

HI

Common strong bases

$\left.\begin{array}{l} NaOH \\ KOH \end{array}\right]$ alkali metal hydroxides (Group IA)

$Ca(OH)_2$

SIMPLE pH CALCULATIONS: STRONG ELECTROLYTES

- With strong acids and bases, the acid or base completely ionizes in water. So, we only have to worry about the effect of the acid or base on the water equilibrium itself.

- Since the equilibrium constant for the self-ionization of water is so small, the strong acid or base will overpower the hydronium (for acids) or hydroxide (for bases) produced by the water.

$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^- \ ; \ K_w = 1.0 \times 10^{-14}$$

Consider a solution of 0.025 M nitric acid (a strong acid):

$$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$$ (Strong acid completely ionizes)

Since the presence of the acid supresses ionization of water (since the acid provides hydronium) and the hydronium produced by water even at neutral pH is much less than 0.025 M, we'll assume that all the hydronium in solution comes from the acid.

$$[H_3O^+] = 0.025\,M$$

What's the pH?

$$pH = -\log_{10}(0.025\,M) = \boxed{1.60}$$

We made an assumption - that the water's own ionization to make hydronium was small enough to ignore. Let's check it! Since water's equilibrium reaction always produces hydronium and hydroxide in a 1:1 ratio, we can see how much hydronium water produces by looking at the HYDROXIDE concentration!

$$[H_3O^+][OH^-] = 1.0 \times 10^{-14} \ ; \ (0.025\,M)[OH^-] = 1.0 \times 10^{-14}$$

$$[OH^-] = 4.0 \times 10^{-13} = [H_3O^+] \text{ made by } H_2O$$

Very, very small compared to 0.025 M !

Consider a solution of 0.0125 M sodium hydroxide (a strong base):

$NaOH \rightarrow Na^+ + OH^-$ (Strong base shuld completely ionize)

The presence of the base should suppress the ionization of water itself, meaning that all the hydroxide ion in solution should come from the base. This is similar to the argument we just made about the acid.

$[OH^-] = 0.0125\ M$ (from the base)

What's the pH?

$pOH = -\log_{10}[OH^-] = 1.90$

$pH + 1.90 = 14.00$ $(pH + pOH = 14.00)$

$pH = \boxed{12.10}$

Like we did before, let us prove that the hydroxide ion from the water itself is small enough to ignore. Since the amount of hydroxide ion produced equals the amount of hydronium produced by the water, ...

$[H_3O^+] = 10^{-pH} = 10^{-12.10} = 7.9 \times 10^{-13}\ M\ H_3O^+$

$= 7.9 \times 10^{-13}\ M\ OH^-$ from water. ✓

Very, very small!