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...

ACIDIC

$[H^+] > 1 \times 10^{-7}$

ALKALINE

$[H^+] < 1 \times 10^{-7}$

pH = 0

pH = 7

NEUTRAL

$[H^+] = 1 \times 10^{-7}$

pH = 14

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 wate 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

$NaOH$, KOH] 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.

$2\,H_2O \rightleftharpoons H_3O^+ + OH^-$

Consider 0.025 M HNO_3

Assume all H_3O^+ comes from acid!

$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$

So, $[H_3O^+]$ = original $[HNO_3] = 0.025$ M

$pH = -\log_{10}[H_3O^+] = -\log_{10}(0.025) = \boxed{1.60}$

For logarithms, the places after the decimal are counted as significant. The places before are merely the exponent.

What would the pOH be?

$pH + pOH = 14.00$

$pOH = 12.40$

... but we usually don't care what pOH is.

$2H_2O \rightleftharpoons H_3O^+ + OH^-$

Consider 0.0125 $\underline{M}$ NaOH

Assume all OH^- comes from base

$$NaOH \rightarrow Na^+ + OH^-$$

$[OH^-]$ = original $[NaOH]$ = 0.0125 $\underline{M}$

$$pOH = -\log_{10}[OH^-] = -\log_{10}(0.0125) = 1.90$$

... now change to pH so we can compare this to the acid problem we just worked

$$pH + pOH = 14.00$$

$$pH + 1.90 = 14.00$$

$$\boxed{pH = 12.10}$$

Let's find the concentration of the hydronium ion, since that will equal the amount of hydroxide produced by the water. (This had better be a lot smaller than the 0.0125 M hydroxide from the base!)

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

... and it IS much smaller than 0.0125!

(A) What is the concentration of hydronium ion in an aqueous solution whose pH is 10.50? (B) What is the hydroxide ion concentration? (C) What molar concentration of sodium hydroxide solution would provide this pH?

A) $pH = 10,50$ $\qquad [H_3O^+] = 10^{-pH} = 10^{-10,50} = \boxed{3,16 \times 10^{-11}\ M\ H_3O^+}$

B) $3,16 \times 10^{-11}\ M = [H_3O^+]$ $\qquad 2H_2O \rightleftarrows H_3O^+ + OH^-$

$K_w = 1,0 \times 10^{-14} = [H_3O^+][OH^-]$

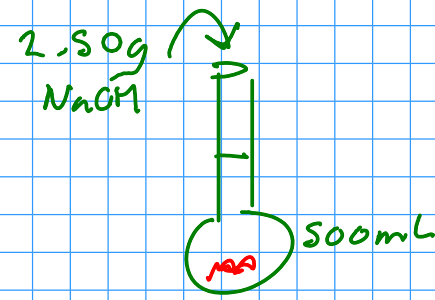
$1,0 \times 10^{-14} = (3,16 \times 10^{-11})[OH^-]$

$\boxed{3,16 \times 10^{-4} = [OH^-]}$

C) $NaOH \longrightarrow Na^+ + OH^-$

1:1, so we need $\boxed{3,16 \times 10^{-4}\ M\ NaOH}$

What is the pH of a sodium hydroxide solution made from dissolving 2.50 g of sodium hydroxide in enough water to make 500.0 mL of solution?



NaOH: 40.00 g/mol

Find concentration of NaOH...

$$2.50\ g\ NaOH \times \frac{mol}{40.00\ g} = 0.0625\ mol\ NaOH$$

$$\underline{M} = \frac{mol}{L} = \frac{0.0625\ mol}{0.5000\ L} = 0.125\ \underline{M}\ NaOH$$

Since NaOH is a strong base, hydroxide concentration equals NaOH concentration ...

$$NaOH \rightarrow Na^+ + OH^-$$

so $[OH^-] = 0.125\ M$

$pOH = 0.90$

$pH = 14.00 - pOH = \boxed{13.10 = pH}$