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 = 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):

$$HNO_3(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + NO_3^-(aq)$$

Assume all of the hydronium ion present at equilibrium comes from the acid (contribution of water itself is minimal)

$$\text{So, } [H_3O^+] = [HNO_3]_{nominal} = 0.025 \, \underline{M} \; H_3O^+$$

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

For logarithms, the places AFTER the decimal point are significant digits, while the numbers in front of the decimal point are not (they;re essentially the exponent of the original number!)

What would the hydroxide ion concentration be?

$$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

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

$$[OH^-] = 4.0 \times 10^{-13}\,M$$

... this is equal to the amount of water that ionizes!

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

Assume all hydroxide ion comes from the base.

$$NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$$

$$[OH^-] = [NaOH]_{nominal} = 0.0125\ M\ OH^-$$

We'd like to know pH. First, calculate pOH.

$$pOH = -\log_{10}(0.0125) = 1.90$$

... but we want pH. pH and pOH are related by the water equilibrium.

$$pH + pOH = 14.00$$

$$pH + 1.90 = 14.00$$

$$\boxed{pH = 12.10}$$

Let's find the concentration of the hydronium ion, since that will equal the amount of water that ionizes. We assumed that the amount of water that ionozes was much less than the amount of base we added to make the solution.

$$[H_3O^+] = 10^{-pH} = 10^{-12.10}$$

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

... This is a much smaller number 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) $pH = 10.50$, $[H_3O^+] = ?$

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

$[H_3O^+] = 10^{-10.50} = \boxed{3.2\times10^{-11}\ \underline{M}\ H_3O^+}$

B) $[H_3O^+][OH^-] = 1.0\times10^{-14}$

$(3.2\times10^{-11})[OH^-] = 1.0\times10^{-14}$

$\boxed{[OH^-] = 3.2\times10^{-4}\ M}$

$pH + pOH = 14.00$

$pOH = 14.00 - 10.50$

$pOH = 3.50$

$[OH^-] = 10^{-3.50} = 3.2\times10^{-4}\ M$

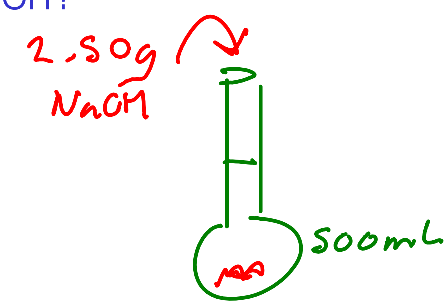
C) Sodium hydroxide is a STRONG BASE, so

$NaOH \rightarrow Na^+ + OH^-$ | 1:1 ratio of $NaOH : OH^-$

$[NaOH]_{nominal} = 3.2\times10^{-4}\ \underline{M}$

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 molarity of NaOH solution

$$M = \frac{\text{mol NaOH}}{\text{L solution}} \leftarrow 0.500\text{ L}$$

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

$$M = \frac{\text{mol NaOH}}{\text{L solution}} = \frac{0.0625\text{ mol NaOH}}{0.500\text{ L}} = 0.125\text{ M NaOH}$$

Sodium hydroxide is a STRONG BASE, so we expect it to completely ionize in solution and control the amount of hydroxide ion present.

$$NaOH \rightarrow Na^+ + OH^- \;;\; [OH^-] = [NaOH]_{nominal} = 0.125\text{ M } OH^-$$

$$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

$$[H_3O^+](0.125) = 1.0 \times 10^{-14}$$

$$[H_3O^+] = 8.0 \times 10^{-14}$$

$$\boxed{pH = 13.10}$$

$$pOH = -\log_{10}(0.125) = 0.90$$

$$pH + pOH = 14.00$$

$$pH = 14.00 - 0.90$$

$$pH = 13.10$$