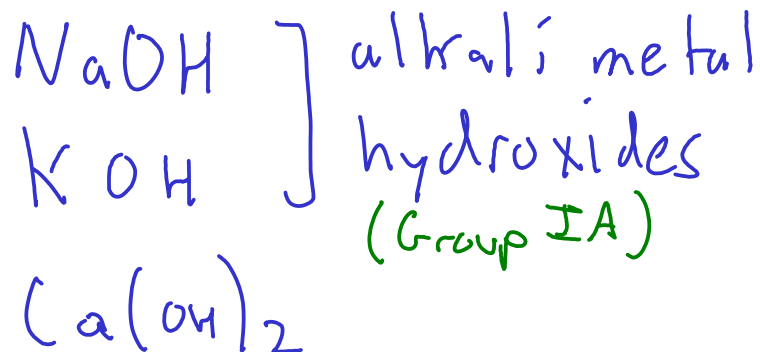# ACID-BASE EQUILIBRIUM IN WATER

- 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

$NaOH$ ⎤ alkali metal
KOH ⎦ 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.

$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^- \; ; \; K_w = 1.0 \times 10^{-14}$$
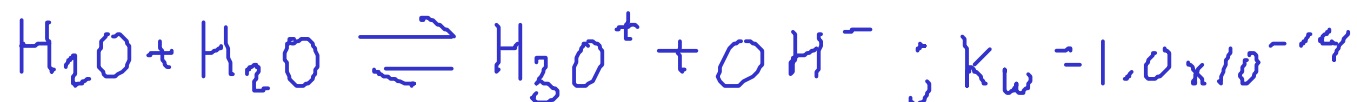
Consider a solution of 0.025 M nitric acid (a strong acid):

$$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$$ (Strong acid completely ionizes!)
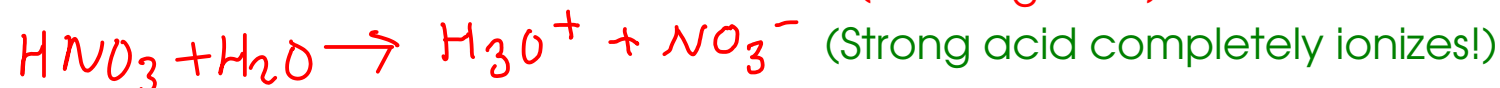
Since the presence of the strong acid suppresses water's self-ionization (the acid makes hydronium ion, a product of water's ionization), we expect that the amount of hydronium ion produced by water itself will be insignificant compared to the hydronium produced by the acid!

$$[H_3O^+] = [HNO_3]_{nominal} = 0.025 \, m \; H_3O^+$$

What's the pH?

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

We made the assumption that water's own ionization produced very little hydronium ion in the presence of the strong acid. We can check this by calculating the HYDROXIDE concentration, since any hydroxide would have come from the water equilibrium and water makes a hydronium ion every time it makes a hydroxide ion.

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

$$[OH^-] = 4.0 \times 10^{-13} \, m = [H_3O^+] \; from \; H_2O$$

Compare! --> $0.025 \, m$ vs (acid)
$0.00000000000004 \, m$ (water)

We've proven that water's hydronium concentration contribution is ignorable!

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

$$NaOH \longrightarrow Na^+ + OH^- \text{ (Strong bases ionize completely!)}$$

Similar to the acid, we'll assume that the presence of a strong base will suppress water's ionization. This means that the HYDROXIDE ion concentration should be set by the strong base. $[OH^-] = [NaOH]_{nominal} = 0.0125 \, M \, OH^-$

What's the pH?

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

$$pH + pOH = 14.00; \quad pH + 1.90 = 14.00 = \boxed{12.10}$$

Like we did before, let's prove that the contribution of the water to the hydroxide concentration is ignorable! Let's find out the hydronium ion concentration, since it will equal the hydroxide concentration produced by the water.

$$[H_3O^+] = 10^{-pH} = 10^{-12.10} = 8.0 \times 10^{-13} \, M \, H_3O^+$$

$$= 8.0 \times 10^{-13} \, M \, OH^- \text{ from } H_2O$$

We've shown that the hydroxide concentration from the water itself is ignorable compared to the hydroxide made by the base!
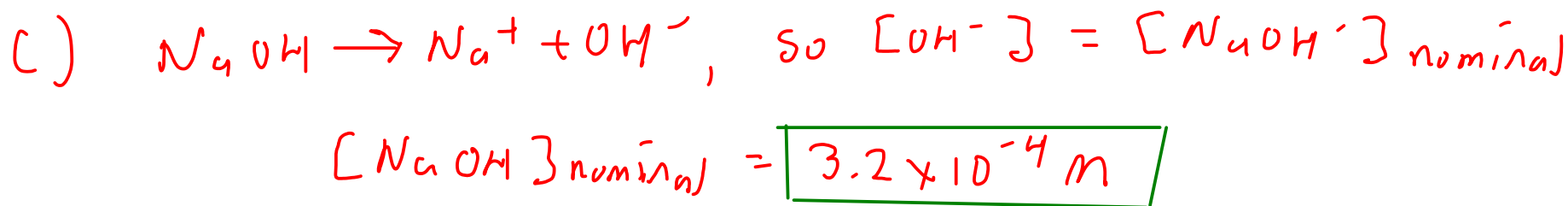
(A) What is the concentration of hydronium ion in an aqueous solution whose pH is 10.50? (B) What is the hydroxide ion concentration? (C) What molar concentration of sodium hydroxide solution would provide this pH?

A) $[H_3O^+] = 10^{-10.50}$ $\left([H_3O^+] = 10^{-PH}\right)$

$= \boxed{3.2 \times 10^{-11} \, M \, H_3O^+}$ (Less than 1.0e-7 M of distilled water)

B) $(3.2 \times 10^{-11})[OH-] = 1.0 \times 10^{-14}$ $\left([H_3O^+][OH^-] = 1.0 \times 10^{-14}\right)$

$[OH^-] = \boxed{3.2 \times 10^{-4} \, M \, OH^-}$

C) $NaOH \longrightarrow Na^+ + OH^-$, so $[OH^-] = [NaOH]_{nominal}$

$[NaOH]_{nominal} = \boxed{3.2 \times 10^{-4} \, M}$

$pH + pOH = 14.00$ $pH = -\log_{10}[H_3O^+]$ $pOH = -\log_{10}[OH^-]$
$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$ $[H_3O^+] = 10^{-pH}$ $[OH^-] = 10^{-pOH}$

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

This is a strong base, so the hydroxide made by the strong base will control the solution's hydroxide concentration. First find out the concentration of NaOH.
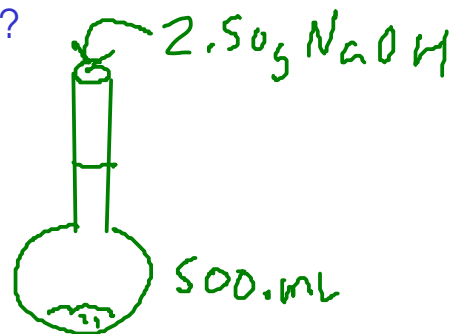
2.50 g NaOH $\times$ $\dfrac{mol\ NaOH}{40.00\ g\ NaOH}$ = 0.0625 mol NaOH

$M = \dfrac{mol\ NaOH}{L\ solution} = \dfrac{0.0625\ mol\ NaOH}{0.5000\ L}$ = 0.125 m NaOH

(500.0 mL)

Since NaOH is a strong base ...

NaOH $\rightarrow$ Na$^+$ + OH$^-$, so $[OH^-] = [NaOH]_{nominal}$

$[OH^-] = 0.125$ m,   pOH $= -\log_{10}(0.125)$ pOH $= -\log_{10}[OH^-]$

$= 0.90$

pH + 0.90 = 14.00

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