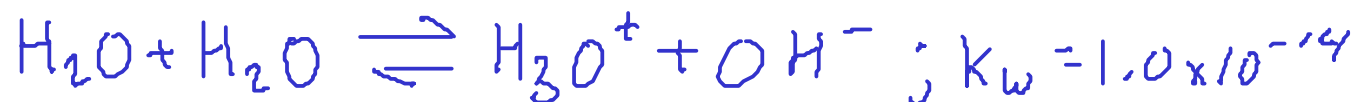


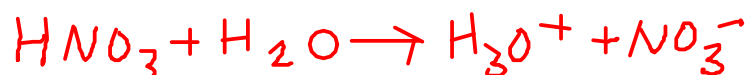
SIMPLE pH CALCULATIONS: STRONG ELECTROLYTES

- 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.



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



Since the water itself only produces a small amount of hydronium ion, and the presence of hydronium ion from the acid means that water's equilibrium reaction will be suppressed (Le Chateleur's Principle), we can assume the concentration of hydronium ion in the solution is equal to the hydronium concentration produced by the acid.

$$[\text{H}_3\text{O}^+] = [\text{HNO}_3]_{\text{nominal}} = 0.025 \text{ M}$$

$$\text{pH} = -\log_{10} [\text{H}_3\text{O}^+] = 1.60^*$$

*Note: For pH, significant digits are the ones to the right of the decimal. The numbers on the left represent the exponent of the original number.

We assumed that the water's ionization was suppressed by the addition of the nitric acid. Let's calculate water's ionization under these conditions. We can do that by calculating HYDROXIDE concentration, since ONLY water produces hydroxide here...

$$\begin{aligned} [\text{H}_3\text{O}^+][\text{OH}^-] &= 1.0 \times 10^{-14} \\ (0.025)[\text{OH}^-] &= 1.0 \times 10^{-14} \\ [\text{OH}^-] &= 4.0 \times 10^{-13} \text{ M} \end{aligned}$$

This number also equals the hydronium concentration produced by the water ... notice it's FAR smaller than 0.025 M!

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



We're going to assume that the hydroxide ion produced by NaOH suppresses water's own ionization - and that the hydroxide concentration will simply equal the nominal NaOH concentration. This is similar to what we said about hydronium in the strong acid example.

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

If we want to know pH, we'll have to find out something about HYDRONIUM ion!

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

$$\text{pOH} = 1.90$$

$$\text{since } \text{pH} + \text{pOH} = 14, \quad \text{pH} = \underline{12.10}$$

Let's check our assumption. Calculate HYDRONIUM concentration, since hydronium is here ONLY made by water!

$$10^{-\text{pH}} = [\text{H}_3\text{O}^+] \quad 10^{-12.10} = 8.0 \times 10^{-13} \text{ M} = [\text{H}_3\text{O}^+]$$

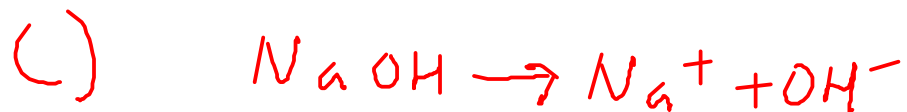
This also equals the amount of HYDROXIDE produced directly by WATER, and (like the hydronium produced by water in the last example) it's so small compared to the production by NaOH that we can ignore it.

(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?

$$A) \quad [H_3O^+] = 10^{-pH} = 10^{-10.50} = 3.2 \times 10^{-11} \text{ M } H_3O^+$$

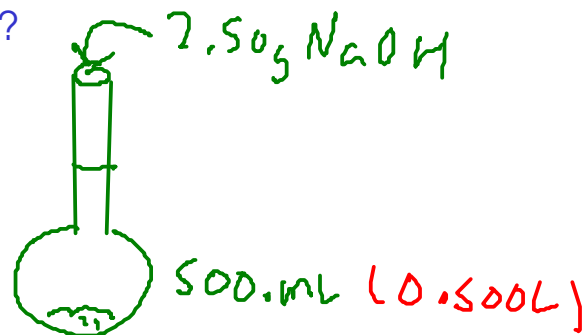
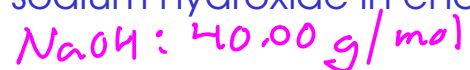
$$B) \quad [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$
$$(3.2 \times 10^{-11}) [OH^-] = 1.0 \times 10^{-14}$$

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



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

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?



First, let's find the molarity of the solution!

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

$$M = \frac{\text{mol}}{\text{L}} = \frac{0.0625 \text{ mol NaOH}}{0.500 \text{ L}} = 0.125 \text{ M NaOH} = [\text{OH}^-]$$

Now use pH identities...

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

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

$$\boxed{\text{pH} = 13.10}$$