

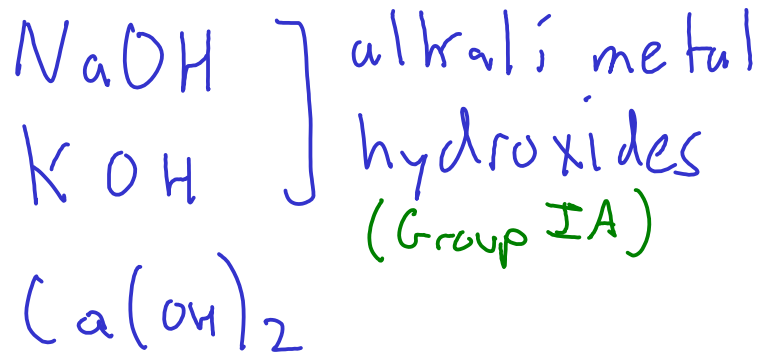
ACID-BASE EQUILIBRIUM IN WATER

- 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

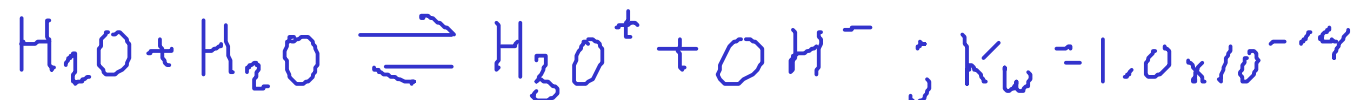


Common strong bases



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):



Assume that all of the hydronium in solution will come from the dissociation of the acid, since the presence of the acid should suppress the ionization of the water.

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

$$\text{pH} = -\log_{10}(0.025) = \boxed{1.60}^*$$

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

What would be the HYDROXIDE ion concentration under these conditions?

$$[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

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

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

<-- This equals the amount of water that has self-ionized in the presence of the acid!

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



Just as before, we will assume that all the HYDROXIDE ION in solution comes from the dissolving of sodium hydroxide.

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

We want to know pH, so we need to change from HYDROXIDE to HYDRONIUM. There are several ways to do that. For this problem, let's find pOH first.

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

... and pOH and pH are related by a very simple equation

$$\text{pH} + \text{pOH} = 14.00 \text{ (@ } 25^\circ\text{C)}$$

$$\text{pH} + 1.90 = 14.00$$

$$\boxed{\text{pH} = 12.10}$$

Let's find out the actual hydronium ion concentration, since that will indicate how much of the water has self-ionized. (We assumed that the amount of water that self-ionized was insignificant!)

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-12.10} \\ = 7.9 \times 10^{-13} \text{ M H}_3\text{O}^+$$

This is much smaller than 0.0125 M...

(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) $\text{pH} = 10.50$, $[\text{H}_3\text{O}^+]$

$$10^{-\text{pH}} = [\text{H}_3\text{O}^+] = 10^{-10.50} = \boxed{3.2 \times 10^{-11} \text{ M H}_3\text{O}^+}$$

B) $[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$

$$(3.2 \times 10^{-11})[\text{OH}^-] = 1.0 \times 10^{-14}$$

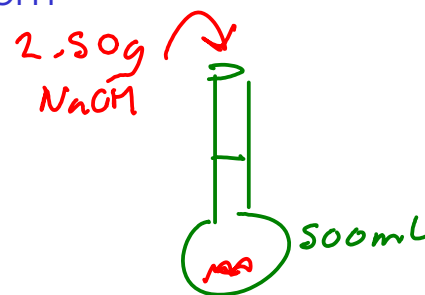
$$[\text{OH}^-] = \boxed{3.2 \times 10^{-4} \text{ M OH}^-}$$



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

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?

$\text{NaOH}: 40.00 \text{ g/mol}$



First, find the molarity of the solution that's described in the problem:

$$M = \frac{\text{mol NaOH}}{\text{L solution}} \leftarrow 0.5000 \text{ L}$$

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

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

Since sodium hydroxide is a STRONG BASE, it will ionize completely. All the hydroxide ion in this solution should come from the sodium hydroxide (it'll suppress the water ionization):



$$[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

$$[\text{H}_3\text{O}^+](0.125) = 1.0 \times 10^{-14}$$

$$[\text{H}_3\text{O}^+] = 8.0 \times 10^{-14} \text{ M}$$

$$\text{pH} = 13.10$$