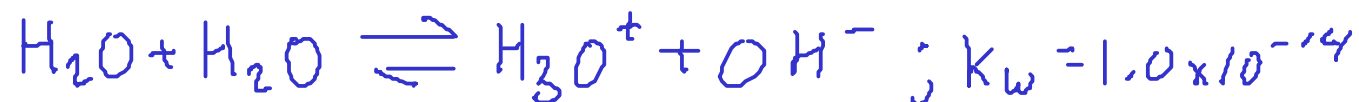


SIMPLE pH CALCULATIONS: STRONG ELECTROLYTES

- 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 the hydronium ion in solution is produced by the nitric acid, since the presence of the acid should suppress water's own production of hydronium (Le Chateleur's Principle)

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

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

* For logarithms, the significant digits are the digits BEHIND the decimal point. The digits in front of the decimal represent the EXPONENT in the original number and are not significant...

What would the HYDROXIDE ion concentration be under these conditions?

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

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

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

... since this also equals the hydronium concentration produced BY WATER ITSELF, we can see that water's production of hydronium really is small enough to be ignored compared to the acid!

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



Similar to the previous example, we will assume that all of the HYDROXIDE ion comes from the dissociation of sodium hydroxide and that the HYDROXIDE will suppress the self-ionization of water...

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

We want to know pH ... How? First find pOH ...

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

... and then use the relationship between pH and pOH to get the answer!

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

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

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

Let's check the hydronium ion concentration, since that will tell us how much water self-ionizes in the presence of the sodium hydroxide ...

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

Since this number also equals the hydroxide ion PRODUCED BY WATER ITSELF, we can see that (like the previous example), the amount is so small we can ignore it compared to the hydroxide produced by the base (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) $\text{pH} = 10.50$, $[\text{H}_3\text{O}^+] = ?$

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

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

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

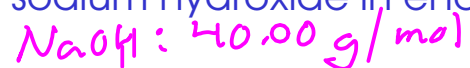
$$[\text{OH}^-] = \boxed{3.2 \times 10^{-4} \text{ M OH}^-}$$



So, 1:1 ratio of NaOH to OH^-

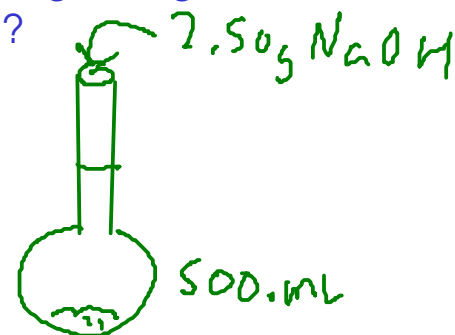
$$[\text{NaOH}]_{\text{nominal}} = \boxed{3.2 \times 10^{-4} \text{ M NaOH}} \quad (0.00032 \text{ M})$$

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?



We will need to know the MOLAR concentration of NaOH, since pH calculations are based on molarity units...

$$M = \frac{\text{mol NaOH}}{\text{L solution}} \leftarrow 500.0 \text{ mL} = 0.5000 \text{ L}$$



To find the molar concentration of NaOH, we'll need to find the moles NaOH ...

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

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

Since NaOH is a strong base, it will completely ionize and set the HYDROXIDE concentration:



$$[\text{OH}^-] = 0.125 \text{ M} \quad \text{Now we need to find pH ...}$$

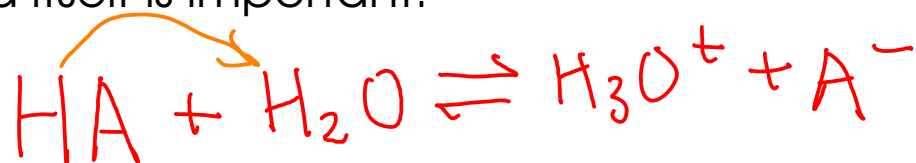
$$[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

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

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

$$\text{pH} = 13.10$$

For a WEAK ACID, equilibrium does not lie far to the right. The ionization equilibrium of the acid itself is important!



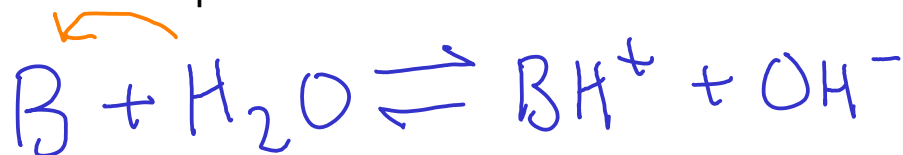
$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

acid ionization constant

Again, water's concentration will not change significantly, so it is folded into the ionization constant

(HA) = concentration of undissociated acid

For a WEAK BASE, equilibrium does not lie far to the right. The ionization equilibrium of the base itself is important!



$$K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]}$$

base ionization constant

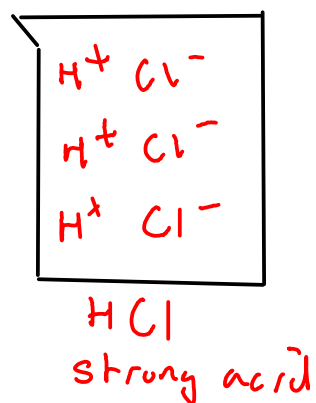
Values for K_a and K_b can often be found in data books / tables / or on the web.

In Ebbing, this data is in the appendices, on pages A-13 and A-14

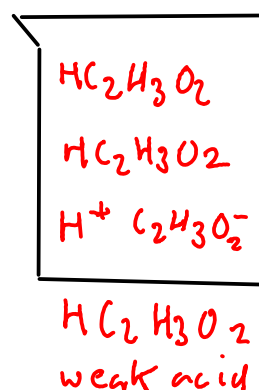
WEAK ELECTROLYTES

- In solutions of weak acids or bases, the UNDISSOCIATED form is present in significantly high concentration.

- The pH of a solution of weak acid will be HIGHER than the pH of a strong acid solution with the same nominal concentration!



VS



Fewer molecules of the weak acid ionize, so the concentration of hydrogen/hydronium ion is lower, meaning a higher pH!

- The pH of a solution of weak base will be LOWER than the pH of a strong base solution with the same nominal concentration!