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

$NaOH \rightarrow Na^+ + OH^-$ (strong base)

So, $[OH^-] = [NaOH]_{nominal}$

So to find the hydroxide concentration, we'll first need to calculate NaOH concentration...

Find molarity ... $\frac{\text{mol NaOH}}{\text{L solution}}$ ← 500.0 mL = 0.5000 L

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

$$\frac{0.0625 \text{ mol NaOH}}{0.5000 \text{ L}} = 0.125 \text{ M NaOH} = 0.125 \text{ M } OH^-$$

$$pOH = -\log(0.125) = 0.90$$

$(pOH = -\log[OH^-])$

$$pH = 14.00 - 0.90 = \boxed{13.10}$$

$(pH + pOH = 14.00)$

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

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

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!

$$B + H_2O \rightleftharpoons BH^+ + OH^-$$

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

base ionization constant

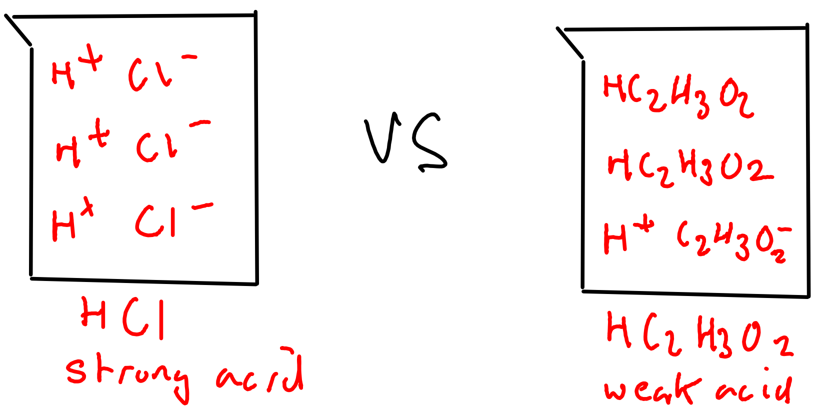
Values for Ka and Kb can often be found in data books / tables / or on the web.

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!



Fewer molecules of the weak acid ionize, so the concnetration 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!

Consider a 0.100M solution of nitrous acid, a WEAK ACID (HNO_2)

$$HNO_2 + H_2O \rightleftharpoons H_3O^+ + NO_2^-$$

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = 4.5 \times 10^{-4}$$

Found on page A-14 in Ebbing 10th edition. These K values are determined experimentally like other equlibrium constants.

What is the pH of the solution?

Unlike the strong acid, we can't assume that all the acid ionizes. Instead, we will actually have to solve the equilibrium problem for nitric acid's ionization!

Species	[Initial]	Δ	[Equilibrium]
H_3O^+	0	$+X$	X
NO_2^-	0	$+X$	X
HNO_2	0.100	$-X$	$0.100 - X$

Let "x" equal the increase in hydronium ion concentration!

$$\frac{(x)(x)}{(0.100-x)} = 4.5 \times 10^{-4}$$

$$\frac{x^2}{0.100-x} = 4.5 \times 10^{-4}$$

Very similar to Chapter 14!

$$\frac{x^2}{0.100 - x} = 4.5 \times 10^{-4}$$

This is a quadratic, We can solve it with the quadratic equation:

$$ax^2 + bx + c = 0$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

We can solve this with the quadratic equation, but ...

1) Assume "x" is small. If it is ...

2) 0.100 - x = 0.100

How do we know whether "x" really is small relative to the starting concentration? Compare Ka and the initial concentration. If they differ by a factor of 1000 or more, the it's generally safe to assume x is small enough to ignore in the subtraction...

$$\frac{x^2}{0.100} = 4.5 \times 10^{-4}$$

$$x = 0.0067082039 = [H_3O^+]$$

$$pH = 2.17$$

($pH = 2.19$ by the quadratic equation)

Compare:

- Weak acid HNO_2 : pH of 0.10 M solution = 2.17

Let's compare the pH of the weak nitrous acid with the pH of a strong acid like nitric acid:

0.10 M HNO_3, what is pH?

$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$

0.10 M HNO_3, $[H_3O^+] = 0.10$ M

pH = 1.00

The stronger the acid:

- the lower the pH of a solution of given concentration will be
- the higher the concentration of hydronium ion (when compared to the nominal acid concentration)

Consider an 0.100 M solution of the weak base ammonia:

NH_3 ; $K_b = 1.8 \times 10^{-5}$ (p A-14, Ebbing 9th)

What is the pH?

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = 1.8 \times 10^{-5}$$

We'll need to find out the HYDROXIDE ion concentration. Then we can convert to hydronium.

Species	[Initial]	Δ	[Equilibrium]
NH_4^+	0	+x	x
OH^-	0	+x	x
NH_3	0.100	−x	0.100−x

Let "x" equal the increase in ammonium ion concentration

$pOH = -\log[OH^-]$

$pH + pOH = 14.00$

$$\frac{(x)(x)}{(0.100-x)} = 1.8 \times 10^{-5}$$

$$\frac{x^2}{0.100-x} = 1.8 \times 10^{-5}$$

1) Assume x is much smaller than 0.100
2) 0.100 - x = 0.100

$$\frac{x^2}{0.100} = 1.8 \times 10^{-5}$$

$x = 0.0013416408 = [OH^-]$

$pOH = 2.87$

$pH = 14.00 - 2.87 = 11.13$

(Solving the quadratic for this problem gives us a pH of 11.13 ... same as with the assumption!)

Compare pH to the pH of an 0.100 M solution of the strong base NaOH:

$pH_{NH_3} = 11.13$

$NaOH \rightarrow Na^+ + OH^-$

So, 0.100 M NaOH, $[OH^-] = 0.100$

$pOH = -\log_{10}(0.100) = 1.00$

$pH = 14.00 - 1.00 = \boxed{13.00}$

The stronger the base:

- the higher the pH will be for a solution of given concentration
- the higher the HYDROXIDE concentration (compared to the nominal base concentration)