

Compare:

- Weak acid HNO_2 : pH of 0.10 M solution = 2.15

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?



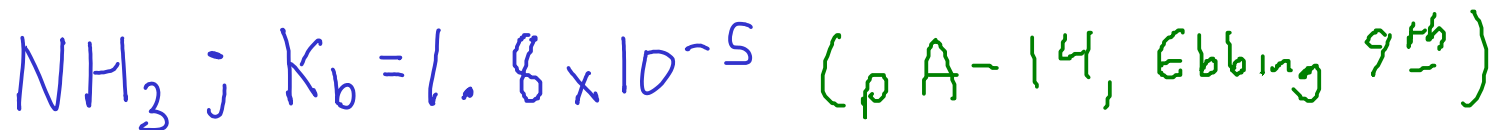
0.10 M HNO_3 , $[\text{H}_3\text{O}^+] = 0.10 \text{ M}$

$$\text{pH} = -\log_{10}(0.10) = 1.00$$

The stronger the acid:

- the lower the pH of a solution of given concentration will be
- the higher the concentration of hydronium ion

140 Consider an 0.100 M solution of the weak base ammonia:



What is the pH?



$$K_b = 1.8 \times 10^{-5} = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

We need to solve this, but which term are we most interested in?

We want to solve for the hydroxide concentration, because it can be converted to hydronium concentration (and pH)!

SPECIES	INITIAL CONC	CHANGE	EQUILIBRIUM CONC
NH_4^+	0	+ X	X
OH^-	0	+ X	X
NH_3	0.100	- X	0.100 - X

Plug in to the equilibrium expression:

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

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

This is a QUADRATIC EQUATION. We can solve it with the quadratic formula, OR we can notice that the value of 'x' should be much smaller than 0.100!

Assume $x \ll 0.100$, so $0.100 - x \approx 0.100$

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

$x = 0.0013416408 = [\text{OH}^-]$ This is HYDROXIDE ion concentration, *NOT* hydronium!

$$\text{pOH} = -\log_{10}(0.0013416408) = 2.87$$

$$\text{pH} = 14.00 - 2.87$$

$$\text{pH} = 11.13$$

If you had solved this problem with the quadratic equation, you would have obtained a pH of 11.13 - no difference at all to two significant figures!

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

$$\text{pH}_{\text{NH}_3} \approx 11.13$$



$$0.100 \text{ M NaOH} ; [\text{OH}^-] = 0.100$$

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

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

The stronger the base:

- the higher the pH will be for a solution of given concentration
- the higher the HYDROXIDE concentration
- the lower the HYDRONIUM concentration

Find the pH and the degree of ionization for an 0.10 M solution of formic acid: HCHO_2



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CHO}_2^-]}{[\text{HCHO}_2]} = 1.7 \times 10^{-4}$$

Constant's value at 25C
obtained from the chart
on p A-13, Ebbing 9th ed

SPECIES	INITIAL CONC	CHANGE	EQUILIBRIUM CONC
H_3O^+	0	+ X	X
CHO_2^-	0	+ X	X
HCHO_2	0.10	- X	0.10 - X

$$1.7 \times 10^{-4} = \frac{(x)(x)}{0.10 - x}$$

$$1.7 \times 10^{-4} = \frac{x^2}{0.10 - x}$$

$$1.7 \times 10^{-4} = \frac{x^2}{0.10 - x}$$

$$\downarrow x \ll 0.10$$

$$1.7 \times 10^{-4} = \frac{x^2}{0.10}$$

$$x = 0.0041231056 = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log_{10}(0.0041231056) = 2.38 = \text{pH}$$

Degree of ionization? DEGREE OF IONIZATION is the fraction of a weak electrolyte (acid or base) that ionizes in water.

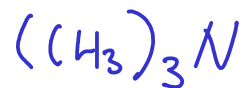
$$\frac{[\text{CHO}_2^-]}{[\text{HCHO}_2]_{\text{orig}}} = \frac{[\text{H}_3\text{O}^+]}{[\text{HCHO}_2]_{\text{orig}}} = \frac{0.0041231056}{0.10} = 0.041 = \text{D.O.I.}$$

Sometimes, we express degree of ionization as a percent and call it PERCENT IONIZATION:

$$\% = \text{D.O.I.} \times 100\% = 4.1\% \text{ ionized}$$

Check this in lab - expt 16A: A more dilute acid solution should have a HIGHER degree of ionization thanks to Le Chatelier's Principle - even if the dilution causes the pH of the acid solution to decrease overall!

An aqueous solution of 0.25 M trimethylamine has a pH of 11.63. What's the experimental value of K_b ?



$$K_b = \frac{[(\text{CH}_3)_3\text{NH}^+][\text{OH}^-]}{[(\text{CH}_3)_3\text{N}]} = \underline{\underline{???}}$$

SPECIES	INITIAL CONC	CHANGE	EQUILIBRIUM CONC
$(\text{CH}_3)_3\text{NH}^+$	0	+ x	x
OH^-	0	+ x	x
$(\text{CH}_3)_3\text{N}$	0.25	- x	0.25 - x

$$K_b = \frac{(x)(x)}{(0.25 - x)}$$

$$K_b = \frac{x^2}{0.25 - x}$$

If we want to find the value of K_b , then we need to come up with some other way (than solving the quadratic equation) of finding 'x'.

$$K_b = \frac{x^2}{0.25 - x}$$

We know from our setup that 'x' equals the hydroxide ion concentration. Since concentration of hydroxide is related to pH, we can find 'x' through the pH

$$x = [\text{OH}^-]$$

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

$$11.63 + \text{pOH} = 14.00$$

$$\text{pOH} = 2.37$$

$$x = 10^{-2.37} = 0.0042657952$$

Now, plug 'x' back into the K_b expression:

$$K_b = \frac{x^2}{(0.25 - x)} = \frac{(0.0042657952)^2}{0.25 - 0.0042657952} = 7.4 \times 10^{-5} = K_b$$