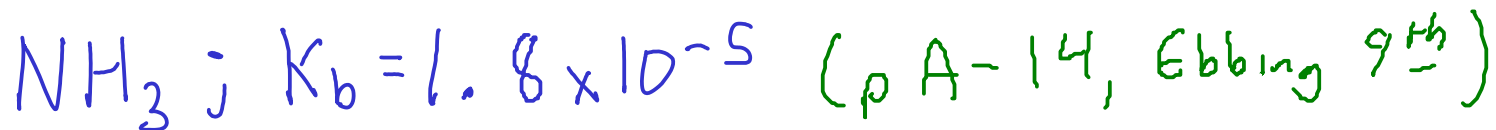
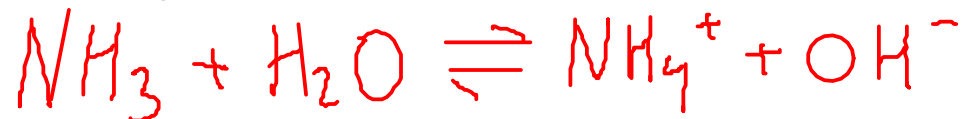


Consider an 0.100 M solution of the weak base ammonia:



What is the pH?



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

We need to solve this, but which one of these terms are we interested in?

We want to solve for the HYDROXIDE concentration, since it's closely related to - and can be easily converted to - pH.

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

Plug into the equilibrium expression:

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

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

This is a QUADRATIC equation. We could solve it with the quadratic formula, BUT like the acid example, we expect that 'x' is small compared to 0.100.

$$0.100 - x \approx 0.100$$

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

$$x = 0.0013416408 = [\text{OH}^-]$$

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

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

$$\boxed{\text{pH} = 11.13}^*$$

Be careful! This is HYDROXIDE concentration, not HYDRONIUM concentration!

* If you had solved this problem using the quadratic, you would have found the pH of the ammonia solution was 11.13 - no difference to two significant figures.

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

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



So, 0.100 M NaOH has $[\text{OH}^-] = 0.100$

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

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

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)

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 was obtained from the chart in Ebbing on page A-13 (9th edition)

Species	[Initial]	Δ	[Equilibrium]
H_3O^+	0	+X	X
CHO_2^-	0	+X	X
HCHO_2	0.10	-X	0.10 - X

$$\frac{(x)(x)}{(0.1-x)} = 1.7 \times 10^{-4}$$

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

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

$$x \ll 0.10$$

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

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

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

$$\text{pH} = 2.38$$

DEGREE OF IONIZATION is the fraction of a weak acid or base that ionizes in water.

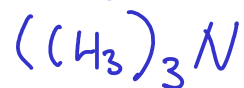
$$\frac{[\text{CHO}_2^-]}{[\text{HCHO}_2]_{\text{initial}}} = \frac{[\text{H}_3\text{O}^+]}{[\text{HCHO}_2]_{\text{initial}}} = \frac{0.0041231056}{0.10} = 0.041 = \text{D.O.I.}$$

Sometimes, we express degree of ionization as a percentage - the PERCENT IONIZATION:

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

Check this in experiment 16A: A more dilute acid solution should have a HIGHER degree of ionization than a more concentrated one thanks to Le Chateleur's principle - even if the dilution causes the pH to increase overall.

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



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

Species	[Initial]	Δ	[Equilibrium]
$(\text{CH}_3)_3\text{NH}^+$	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 of finding the value of 'x'

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

We know that 'x' equals the HYDROXIDE concentration. Since hydroxide concentration is related to pH, we can use the pH to find 'x'.

$$pH = 11.63 ; \quad pH + pOH = 14.00$$

$$pOH = 2.37$$

$$[OH^-] = 10^{-2.37} = 0.0042657952$$

$$[OH^-] = x = 0.0042657952$$

Now, plug 'x' into the equilibrium expression:

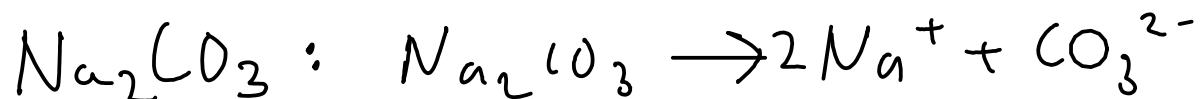
$$K_b = \frac{x^2}{0.25 - x} = \frac{(0.0042657952)^2}{0.25 - 0.0042657952}$$

$$K_b = 7.4 \times 10^{-5}$$

SALTS

- Compounds that result from the reaction of an acid and a base.
- Salts are strong electrolytes (completely dissociate in water) IF SOLUBLE (not all salts dissolve appreciably).
- Most ionic compounds are considered salts (they can be made by some reaction between the appropriate acid and base)
- Salts have acidic and basic properties! The ions that form when salts are dissolved can be acidic, basic, or neutral.

- Salts made from WEAK ACIDS tend to form BASIC solutions
- Salts made from WEAK BASES tend to form ACIDIC solutions



Do any of these ions have acidic or basic properties?

Na^+ : neutral. Not a proton donor or a proton acceptor

CO_3^{2-} : BASIC, since it can accept protons to form the weak acid CARBONIC ACID in solution.



ACID

BASE

SALT OF A WEAK ACID

ex: $\text{NaC}_2\text{H}_3\text{O}_2$ 

For this reaction to occur, HA MUST be stable in water. In other words, a weak acid.



The anion is a BASE. It can accept a proton from water to form the weak (therefore stable as a molecule!) acid HA

$$K_b = \frac{[\text{HA}][\text{OH}^-]}{[\text{A}^-]} \quad \left| \text{--- This is the base ionization constant for } \text{A}^- \right.$$

Since A^- and HA are a conjugate pair, the ionization constants are related!

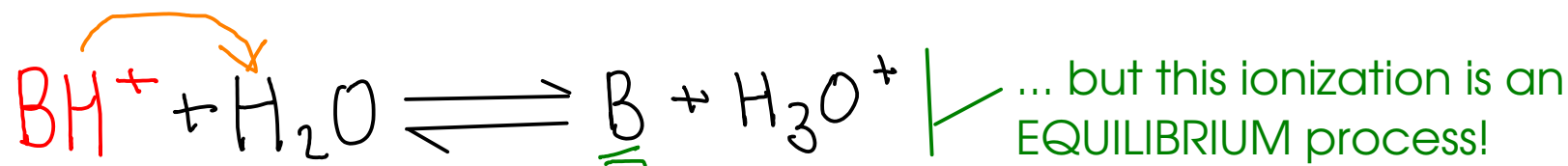
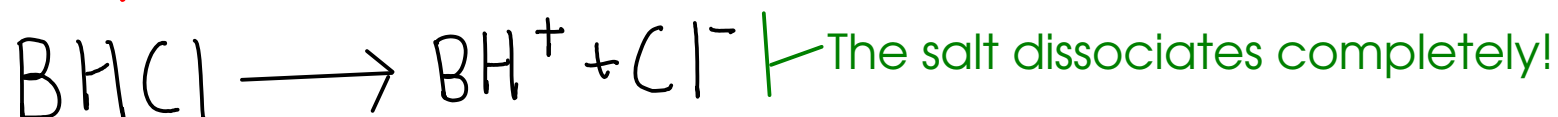
$$K_w = (K_{a,\text{HA}})(K_{b,\text{A}^-})$$

1.0×10^{-14}

$$14 = \text{p}K_a + \text{p}K_b$$

You will generally not find both the K_a AND K_b for a conjugate pair in the literature, since one can be easily converted to the other!

SALT OF A WEAK BASE

ex: NH_4Cl 

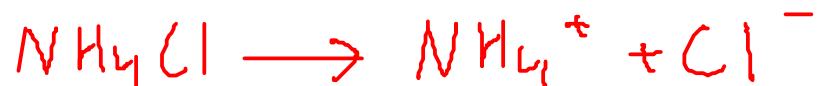
$$K_a = \frac{[\text{B}][\text{H}_3\text{O}^+]}{[\text{BH}^+]} \quad \left| \text{Acid ionization constant for } \text{BH}^+ \right.$$

$$K_w = (K_{a,\text{BH}^+})(K_{b,\text{B}})$$

1.0×10^{-14}

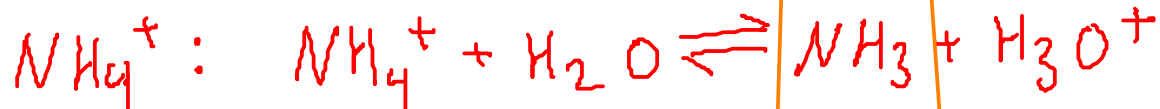
Find the pH for salt solutions just like you would find pH for any other weak acid or weak base solutions. Only trick is to find out whether the salt is actually acidic or basic!

0.100 M NH_4Cl ... Find the pH of the solution



Is this salt acid, basic, or neutral?

This is the WEAK BASE ammonia. Stable in water



This is the STRONG ACID hydrochloric acid, which completely ionizes in water. This is not a stable molecule in water.

The conjugate of a STRONG acid or base is NEUTRAL - it does not affect the pH.

We will need to solve the following equilibrium:



This equilibrium is the only one of the two above that affects the pH!



$$K_{a, \text{NH}_4^+} = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$$

We can get K_a for the acid from K_b of the base (ammonia):

$$K_{b, \text{NH}_3} = 1.8 \times 10^{-5}$$

$$K_a \times K_b = 1.0 \times 10^{-14}$$

$$\text{So, } K_a = 5.56 \times 10^{-10}$$

Species	[Initial]	Δ	[Equilibrium]
NH_3	0	+X	X
H_3O^+	0	+X	X
NH_4^+	0.100	-X	0.100 - X

$$\frac{x^2}{0.100 - x} = 5.56 \times 10^{-10}$$

$x \ll 0.100$

$$\downarrow 0.100 - x \approx 0.100$$

$$\frac{x^2}{0.100} = 5.56 \times 10^{-10}$$

$$x = 7.46 \times 10^{-6}$$

$$\text{So, } [\text{H}_3\text{O}^+] = 7.46 \times 10^{-6}$$

$$\text{pH} = 5.13$$

Compare:

pH = 1.00 for 0.100 M strong acid

pH = 2.16 for 0.100 M nitrous acid (weak acid)

pH = 7.00 for distilled water