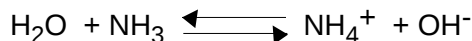


Weak Bases: Just like weak acids, weak bases do not dissociate completely in aqueous solution so we have to look at the equilibrium between the weak base and it's ions. Weak bases are a little harder to think about because most don't give up OH⁻'s, what they do is break apart H₂O to make OH⁻'s. They do this by donating an electron pair to a hydrogen ion thus they demonstrate the Lewis concept of a base.

For example, look at what ammonia does in water:



Now, in general you can look at weak base reactions as:



For lack of a better term, call the unknown weak base B.

Base Dissociation Constant:

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

H₂O is left out of this because we assume its concentration is a constant and just factor it into K_b.

Here are some weak bases

<u>Weak Base</u>	<u>K_b</u>
Ammonia, NH ₃	1.8 × 10 ⁻⁵
Aniline, C ₆ H ₅ NH ₂	4.6 × 10 ⁻¹⁰
Dimethylamine, (CH ₃) ₂ NH	7.4 × 10 ⁻⁴
Hydrazine, N ₂ H ₄	9.8 × 10 ⁻⁷
Methylamine, CH ₃ NH ₂	4.2 × 10 ⁻⁴
Pyridine, C ₅ H ₅ N	1.5 × 10 ⁻⁹
Trimethylamine(CH ₃) ₃ N	7.4 × 10 ⁻⁵

*By the way, K_a and K_b values vary and the values here may differ slightly from those in whatever text you are using.

Use K_b to find [OH⁻] it's exactly the same equilibrium stuff, make the table, see what algebraic assumptions you can make etc.

	$\text{H}_2\text{O} + \text{B} \rightleftharpoons \text{HB}^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$		
	$\frac{[\text{B}]}{[\text{B}]_0}$	$\frac{[\text{HB}^+]}{-}$	$\frac{[\text{OH}^-]}{-}$
initial	[B] ₀	-	-
change	[B] ₀ - x	x	x
final	[B] ₀ - x	x	x

This is assuming that we start by just dumping some B into water

$$K_b = \frac{[\text{HB}^+][\text{OH}^-]}{[\text{B}]} = \frac{x^2}{[\text{B}]_0 - x} \quad \text{or} \quad K_b = \frac{x^2}{[\text{B}]_0}$$

You can make the simplifying assumption if K_b < 10⁻⁵

Things can appear in both the weak acid and weak base tables. Molecules that can function as both acid and base are called amphiprotic. Also if you compare K_a and K_b values you'll notice that:

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

Student Packet: 6.4: Weak Bases: AssessMe questions

- Which of the following sets ranks the Brønsted acids shown from strongest to weakest?
A) HCl, acetic acid, HF, H₂O (acetic acid is CH₃COOH)
B) HF, acetic acid, HCl, H₂O
C) HCl, HF, H₂O, acetic acid
D) HCl, HF, acetic acid, H₂O
E) I'm clueless
- Which is a stronger Brønsted base HSO₄⁻, H₂PO₄²⁻, or F⁻?
A) HSO₄⁻
B) H₂PO₄²⁻
C) F⁻
D) These are not Brønsted bases.
E) I'm clueless
- If the pK_a of the acid HX is 8.0, the K_b for X⁻ is?
A) 10⁸
B) 10⁻⁸
C) 10⁶
D) 10⁻⁶
E) I'm clueless
- Which of the following is a base-conjugate acid pair?
A) C₆H₅NH₂, C₆H₅NH₃⁺
B) Cl⁻, NH₄⁺
C) NH₃, HC₂H₃O₂
D) HC₂H₃O₂, H₂O
E) I'm clueless
- How many milliliters of 0.20 M AlCl₃ solution would be necessary to precipitate all of the Ag⁺ from 45 ml of a 0.20 M AgNO₃ solution?
$$\text{AlCl}_{3(\text{aq})} + 3\text{AgNO}_{3(\text{aq})} \longrightarrow \text{Al}(\text{NO}_3)_{3(\text{aq})} + 3\text{AgCl}_{(\text{s})}$$

A) 15 ml
B) 30 ml
C) 45 ml
D) 60 ml
E) I'm clueless
- Which of the following salts has the greatest molar solubility in pure water?
A) CaCO₃ (K_{sp} = 8.7 × 10⁻⁹)
B) CuS (K_{sp} = 8.5 × 10⁻⁴⁵)
C) Ag₂CO₃ (K_{sp} = 6.2 × 10⁻¹²)
D) Pb(IO₃)₂ (K_{sp} = 2.6 × 10⁻¹³)
E) I'm clueless

Student Packet: 6.4: Weak Bases: ShowMe questions

- 1) Determine the pH of a 0.05 M NH_3 solution ($K_b = 1.8 \times 10^{-5}$).

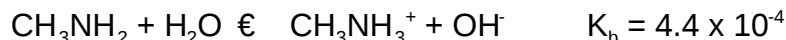
- 2) What is the pH of a 0.18 M solution of NH_4Cl ?
(K_b of $\text{NH}_3 = 1.8 \times 10^{-5}$)

- 3) Determine the percent dissociation of a 2.25 M weak base solution with a pH = 11.8.

- 4) Calculate the pH of a solution prepared by mixing 20.0 mL of 0.05 M HCl and 50.0 mL of 0.40 M $\text{C}_6\text{H}_5\text{NH}_2$. ($K_b = 3.8 \times 10^{-10}$)

Student Packet: 6.4: Weak Bases: TestMe questions

1) Consider the reaction:



1.0 mol of KOH is added to a solution made by adding 2.0 mol CH_3NH_2 plus 1.0 L of H_2O (assume no volume change on addition of solutes). What is the equilibrium $[\text{CH}_3\text{NH}_3^+]$?

- A) 3.2×10^{-2} M
- B) 2.2×10^{-4} M
- C) 2.0×10^{-3} M
- D) 8.8×10^{-4} M
- E) None of the above

2) What is the pH of a 1.0 M solution of aniline? ($K_b = 4 \times 10^{-10}$)

- A) 4.7
- B) 9.3
- C) 9.4
- D) 8.6
- E) None of the above

3) Calculate the pH of a solution made by mixing equal volumes of 0.04 M $\text{C}_5\text{H}_5\text{N}$ ($K_b = 1.7 \times 10^{-9}$) and 0.04 M HNO_3 .

- A) 3.46
- B) 3.66
- C) 5.23
- D) 5.08
- E) None of the above

4) A 0.20 M solution of a weak base is 2.4% ionized. What is the K_b of this base?

- A) 1.5×10^{-2}
- B) 1.2×10^{-4}
- C) 2.5×10^{-2}
- D) 2.5×10^{-4}
- E) None of the above

5) What is the NH_4^+ concentration of a 1.0 L, 0.15 M NH_3 solution in which 1.3 g NaOH has been added (assume no volume change). K_b of NH_3 is 1.8×10^{-5} .

- A) 4.0×10^{-6} M
- B) 6.4×10^{-5} M
- C) 8.2×10^{-5} M
- D) 8.7×10^{-5} M
- E) 3.4×10^{-4} M