

**Simple Step-by-Step
Guides to Solving
Chemistry Problems**

Weak Acids & Bases

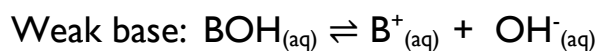
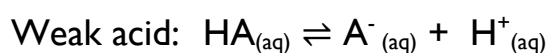


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Weak Acids & Bases

- Calculate pK_a from acid dissociation constant, K_a
- Calculate pK_b from base dissociation constant, K_b
- Calculate $[H^+]$, $[OH^-]$ and pH from acid dissociation constant, K_a
- Calculate $[H^+]$, $[OH^-]$ and pH from base dissociation constant, K_b

Weak acids, H_nA and weak bases, $B(OH)_n$ partially dissociate (ionise) in water:



The degree of dissociation is given by the acid (K_a) and the base (K_b) dissociation constants.

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

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

As the word 'constant' implies, at a given temperature, K_a and K_b . The larger the value of K_a and K_b , the stronger the acid and base, respectively. It is often convenient to express K_a and K_b as ordinary numbers

$$\text{Weak acids: } \quad pK_a = -\log K_a \quad K_a = 10^{-pK_a}$$

$$\text{Weak bases: } \quad pK_b = -\log K_b \quad K_b = 10^{-pK_b}$$

Calculate pK_a from K_a

Essential Equation:

$$pK_a = -\log(K_a)$$

Example 1: What is the pK_a of a weak acid with a K_a of 1.2×10^{-4}

Answer:

$$pK_a = -\log(1.2 \times 10^{-4}) = 3.92$$

Casio Calculator Button Sequence

- **log** **(** **1** **.** **2** **x** **exp** **-** **4** **)** **=**

Calculate pK_b from K_b

Essential Equation:

$$pK_b = -\log(K_b)$$

Example: The K_b value for ammonia is 1.8×10^{-5} . What is the pK_b of ammonia?

$$K_b = -\log(1.8 \times 10^{-5}) = 4.75$$

Casio calculator button sequence

- **log** **(** **1** **.** **8** **x** **exp** **-** **5** **)** **=**

Calculate [H⁺], [OH⁻] and pH from K_a

Essential Equations:

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

$$pH = -\log[H^+]$$

Step 1: Write a balanced equation for the dissociation of the weak acid, ie $HA_{(aq)} \rightleftharpoons A^-_{(aq)} + H^+_{(aq)}$

Step 2: Write an expression for K_a and rearrange to make OH⁻ the subject of the equation and solve.

$$K_a = \frac{[A^-][H^+]}{[HA]} = \frac{[H^+][H^+]}{[HA]} = \frac{[H^+]^2}{[HA]}$$

Note: from the balanced equation for the dissociation, $[A^-] = [H^+]$

Rearranging,

$$[H^+]^2 = K_a \times [HA]$$

$$[H^+] = \sqrt{K_a[HA]}$$

Step 3: Calculate pH and hence pOH, using:

$$pH = -\log_{10} \sqrt{K_a[HA]}$$

$$pOH = 14 - pH$$

Calculate $[\text{OH}^-]$, pOH and pH from K_b

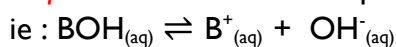
Useful equations:

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

$$\text{pOH} = -\log[\text{OH}^-]$$

$$\text{pH} = 14 - \text{pOH}$$

Step 1: Write a balanced equation for the dissociation of the weak base,



Step 2: Write an expression for K_b and rearrange to make OH^- the subject of the equation and solve.

$$K_b = \frac{[\text{B}^+][\text{OH}^-]}{[\text{BOH}]} = \frac{[\text{OH}^-][\text{OH}^-]}{[\text{BOH}]} = \frac{[\text{OH}^-]^2}{\text{BOH}}$$

Note: from the balanced equation for the dissociation, $[\text{B}^+] = [\text{OH}^-]$

Rearranging,

$$[\text{OH}^-]^2 = K_b \times [\text{BOH}]$$

$$[\text{OH}^-] = \sqrt{K_b}[\text{BOH}]$$

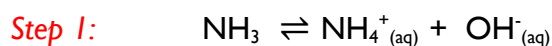
Step 3: Calculate pOH and hence pH, using:

$$\text{pOH} = -\log_{10} \sqrt{K_b}[\text{BOH}]$$

$$\text{pH} = 14 - \text{pOH}$$

Example: What is the pH of a 0.15 M solution of weak base ammonium bromide? The K_b value for ammonia is 1.8×10^{-5} .

Answer:



Step 2: $K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_4\text{OH}]} = \frac{[\text{OH}^-]^2}{[\text{NH}_4\text{OH}]}$

Rearranging, $[\text{OH}^-] = \sqrt{K_b}[\text{NH}_4\text{OH}]$

$$[\text{OH}^-] = \sqrt{(1.8 \times 10^{-5} \times 0.15)} = 1.64 \times 10^{-3}$$

Step 3: Calculate pOH and hence pH, using:

$$\text{pOH} = -\log_{10}(1.64 \times 10^{-3}) = 2.79$$

$$\text{pH} = 14 - 2.79 = \mathbf{11.21}$$

? Practice Problems

- Find the pH of a 0.056 M propionic acid solution ($K_a = 1.4 \times 10^{-5}$).
- Find the pH of a 0.065 M solution of formic acid. The acid dissociation constant (K_a) for formic acid is 1.8×10^{-4} .
- Find the pH of a 0.15 M solution of ammonia, NH_3 . $K_b = 1.8 \times 10^{-5}$.
- Find the pH of a 0.600 M solution of methylamine CH_3NH_2 . $K_b = 4.4 \times 10^{-4}$.
- Calculate $[\text{OH}^-]$ for a 0.50 M solution of ammonia. $K_b = 1.8 \times 10^{-5}$.
- Calculate $[\text{H}^+]$ in a 0.10 M solution of formic acid. $K_a = 1.7 \times 10^{-4}$.
- Determine the value of K_a for acetic acid from the following data: 0.10 mole of the acid is dissolved in enough water for a total volume of 1.0 Litre. The resulting $[\text{H}^+]$ is 1.35×10^{-3} .
- Lactic acid, $\text{CH}_3\text{CH}(\text{OH})\text{COOH}$, gets its name from sour milk, from which it was first isolated in 1780 (L. lactis, milk). K_a for lactic acid is 8.4×10^{-4} . Find the $[\text{H}^+]$ in a sample of sour milk containing 0.100 M lactic acid.
- Many of the common organic acids got their original names from their odors and/or sources. Another case in point is Caproic acid (hexanoic acid), found in the skin secretions of goats (L. caper, goat). Caproic acid is $\text{CH}_3(\text{CH}_2)_4\text{COOH}$ and has a structure similar to acetic acid, but with a longer carbon chain. The concentration of H^+ in a solution prepared by dissolving 0.030 mol of caproic acid in 1.0 L of water solution was measured and found to be 6.5×10^{-4} M. Find K_a for caproic acid.
- Two solutions are needed in the lab, each with a volume of 10 L (to the nearest 1 Litre) and a pH equal to 11.00. How many moles of each solute would it take if one solution is to be made with NaOH (strong base) and the other with NH_3 ? ($K_b = 1.8 \times 10^{-5}$) (approximate for NH_3)

Answers are given in the following page.

? Practice Problem Answers

- a. pH = 3.05
- b. pH = 2.47
- c. pH = 11.22
- d. pH = 12.2
- e. $[\text{OH}^-] = 3 \times 10^{-3}$
- f. $[\text{H}^+] = 4.12 \times 10^{-3}$
- g. 1.8×10^{-5}
- h. $8.8 \times 10^{-3} \text{ M}$
- i. 1.4×10^{-5}
- j. For NaOH, 0.010 moles; for NH₃, 0.56 moles