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

Weak acid: $HA_{(aq)} \rightleftharpoons A^-_{(aq)} + H^+_{(aq)}$

Weak base: $BOH_{(aq)} \rightleftharpoons B^+_{(aq)} + OH^-_{(aq)}$

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

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

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

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 pKb from Kb

Essential Equation:

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

Example: The Kb value for ammonia is 1.8 × 10^{-5}. What is the pKb 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,

$$[H^+]^2 = K_a \text{ x } [HA]$$

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

Step3: Calculate pH and hence pOH, using:

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

$$pOH = 14 - pH$$

Calculate [OH⁻], pOH and pH from K_b

Useful equations:

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

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

$$pH = 14 - pOH$$

Step 1: Write a balanced equation for the dissociation of the weak base, ie : $BOH_{(aq)} \rightleftharpoons B^+_{(aq)} + OH^-_{(aq)}$

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

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

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

Rearranging,

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

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

Step 3: Calculate pOH and hence pH, using:

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

$$pH = 14 - pOH$$

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 1: $NH_3 \rightleftharpoons NH_4^+{}_{(aq)} + OH^-_{(aq)}$

Step 2: $$K_b = \frac{[NH_4^+][OH^-]}{[NH_4OH]} = \frac{[OH^-]^2}{[NH_4OH]}$$

Rearranging, $[OH^-] = \sqrt{K_b[NH_4OH]}$

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

Step 3: Calculate pOH and hence pH, using:

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

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

Practice Problems

a. Find the pH of a 0.056 M propionic acid solution (K_a = 1.4 x 10^{-5}).

b. Find the pH of a 0.065 M solution of formic acid. The acid dissociation constant (K_a) for formic acid is 1.8 x 10^{-4}.

c. Find the pH of a 0.15 M solution of ammonia, NH_3. K_b = 1.8 X 10^{-5}

d. Find the pH of a 0.600 M solution of methylamine CH_3NH_2. K_b = 4.4 x 10^{-4}.

e. Calculate $[OH^-]$ for a 0.50 M solution of ammonia. K_b = 1.8 x 10^{-5}.

f. Calculate $[H^+]$ in a 0.10 M solution of formic acid. K_a = 1.7 x 10^{-4}.

g. Determine the value of Ka 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 $[H^+]$ is 1.35 x 10^{-3}.

h. Lactic acid, $CH_3CHOCOOH$, 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 x 10^{-4}. Find the $[H^+]$ in a sample of sour milk containing 0.100 M lactic acid.

i. 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 $CH_3(CH_2)_4COOH$ 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 x 10^{-4} M. Find K_a for caproic acid.

j. 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? (Kb = 1.8 x 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. $[OH^-] = 3 \times 10^{-3}$
f. $[H^+] = 4.12 \times 10^{-3}$
g. 1.8×10^{-5}
h. 8.8×10^{-3} M
i. 1.4×10^{-5}
j. For NaOH, 0.010 moles; for NH_3, 0.56 moles