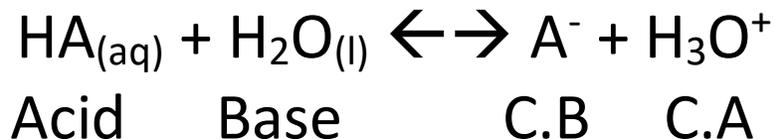


Calculating % Dissociation

In Acid and base questions, we can calculate the amount of ions that are broken or dissociated at equilibrium. Once the dissociated ions are calculated, the % can be calculated by comparing to the original amount.

In other words, we can calculate the amount of ions that were broken apart or dissociated from the original.



In the standard above reaction, the HA is the standard acid where the "H" is the proton that gets donated while the H₂O acts as the base in receiving the extra proton.

1. For example, if a weak acid of HA with a concentration of 0.1M and [H₃O⁺] has a concentration of 0.0005M, we can calculate the amount (in percentage) that HA was dissociated.

$$\frac{0.0005\text{M}}{0.1\text{M}} \times 100 = 0.5\%$$

What does the answer mean?

0.5% of the HA was dissociated into [H₃O⁺] – which isn't too big of an amount, but it makes sense because it is a Weak Acid

Variations

From this question, we can throw in couple curveballs,

Variation 1: Adding in Ka (same for Kb as well)

If Ka is added, the equilibrium concentration can be used and also it is usually a weak acid

1. For example, if a weak acid of HA with a concentration of 0.1M with a $K_a = 1.0 \times 10^{-4}$, calculate the percentage dissociated.

We need to use the ICE table to calculate the $[H_3O^+]$

ICE TABLE:

	HA	+	H ₂ O	→	A ⁻	+	H ₃ O ⁺
I	0.1M		OMIT		0M		0M
C	-x		Because of		+x		+x
E	0.1 - x		LIQUID		+x		+x

$$K_a = \frac{[A^-][H_3O^+]}{[HA]}$$

$$K_a = \frac{[x][x]}{0.1 - x}$$

Assume zero for x
because it is negligible

$$1.0 \times 10^{-4} = \frac{x^2}{0.1}$$

$$x = 0.0032M$$

$$[H_3O^+] = 0.0032M$$

With this value, we can calculate a bunch of things:

- 1) the pH by the pH formula = $-\log[H^+]$
- 2) the $[OH^-]$ by the K_w equation
- 3) the pOH by subtracting the pH from 14 or $-\log[OH^-]$
- 4) % dissociated

Variation 2: When the % dissociated is provided

When the % dissociated is provided, that means the steps must be reversed from the previous example. You now have to use the % provided and calculate the $[H_3O^+]$ dissociated from the total initial concentration. It is like that you have 70% on a test, and the total number of test questions is 40, how many did you score?

For example, Calculate the $[H_3O^+]$ of a 0.88M of a weak acid that is dissociated 0.34%.

Based on the question, you know that only 0.34% of 0.88M was dissociated into $[H_3O^+]$. We can calculate the $[H_3O^+]$ by:

$$\frac{0.34\%}{100} \times 0.88 = 0.00299M \text{ of } [H_3O^+]$$

After calculating for the $[H_3O^+]$, we can use this number to calculate a bunch of things:

- 1) $[OH^-]$ from the K_w equation
- 2) pH from the equation: $pH = -\log[H^+]$
- 3) pOH from the equation: $14 = pH + pOH$
- 4) K_a based on the equilibrium expression since we have the "x" and the original concentration.

Hints: Don't try to memorize the question type – this will only hinder your understanding. Try to relate the meaning of calculating for the $[H_3O^+]$ or $[OH^-]$. Also, understand all the various concentrations that you can calculate if you have pH, pOH, K_w , K_a , K_b , $[H_3O^+]$, and $[OH^-]$ because they all interrelate with each other.