<u>Calculating % Dissociation</u>

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.

$$HA_{(aq)} + H_2O_{(l)} \leftrightarrow A^- + H_3O^+$$

Acid Base C.B C.A

In the standard above reaction, the HA is the standard acid where the "H" is the proton that gets donated while the H_2O 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 $[H_3O^+]$ has a concentration of 0.0005M, we can calculate the amount (in percentage) that HA was dissociated.

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

What does the answer mean?

0.5% of the HA was dissociated into $[H_3O^+]$ – 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 Ka = 1.0×10^{-4}, calculate the percentage dissociated.

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

ICE TABLE:

	HA +	H_2O	→	A^-	+ H_3O^+
I	0.1M	OMIT		0M	0M
C	-x	Because of		+x	+x
E	0.1 - x	LIQUID		+x	+x

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

$$Ka = \frac{[x]\,[x]}{0.1 - x}$$

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 Kw 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:

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

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

1) [OH-] from the Kw equation
2) pH from the equation: pH = -log[H+]
3) pOH from the equation: 14 = pH + pOH
4) Ka 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, Kw, Ka, Kb, $[H_3O^+]$, and [OH-] because they all interrelate with each other.