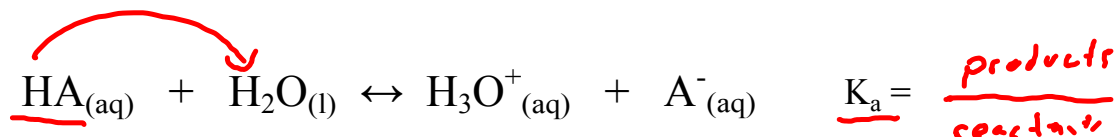


Part 2 - K_a, K_b and pH and pOH calculations

- the acid ionization constant, K_a, is the equilibrium constant for the reaction of an acid with water

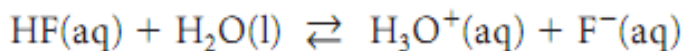


- K_a values are found in the right hand column of your Acid-base chart on pg 8-9 of your data booklet.

- Note that in this table, the first six acids have K_a values given only as “very large.”

- These acids are collectively called the strong acids because they all react quantitatively (>99.9%) with water to form hydronium ions

Example



The equilibrium law expression is

$$K_a = \frac{[\text{H}_3\text{O}^+(\text{aq})][\text{F}^-(\text{aq})]}{[\text{HF}(\text{aq})]} = 6.3 \times 10^{-4}$$

- Two calculations involving the K_a constant are common for weak acid solutions:

1. calculating a K_a value from measured (empirical) amount concentration data
2. using a K_a value to predict a concentration of hydronium ions for an aqueous solution where the initial weak acid amount concentration is known

Calculating K_a from Amount Concentrations

Example 1:

@ equilibrium

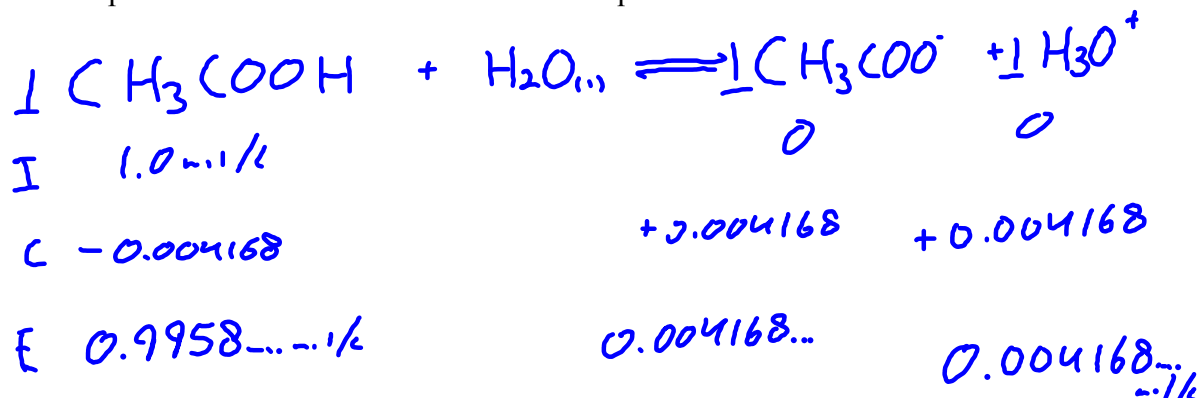
The pH of a 1.00 mol/L solution of acetic acid is carefully measured to be 2.38 at SATP. What is the value of K_a for acetic acid?

Step 1: Write the reaction of the acid with water and write the K_a equilibrium law expression



Step 2: find the equilibrium concentration of aqueous hydronium ion from the pH

Step 3: Use an ICE table to calculate the equilibrium concentration of the weak acid



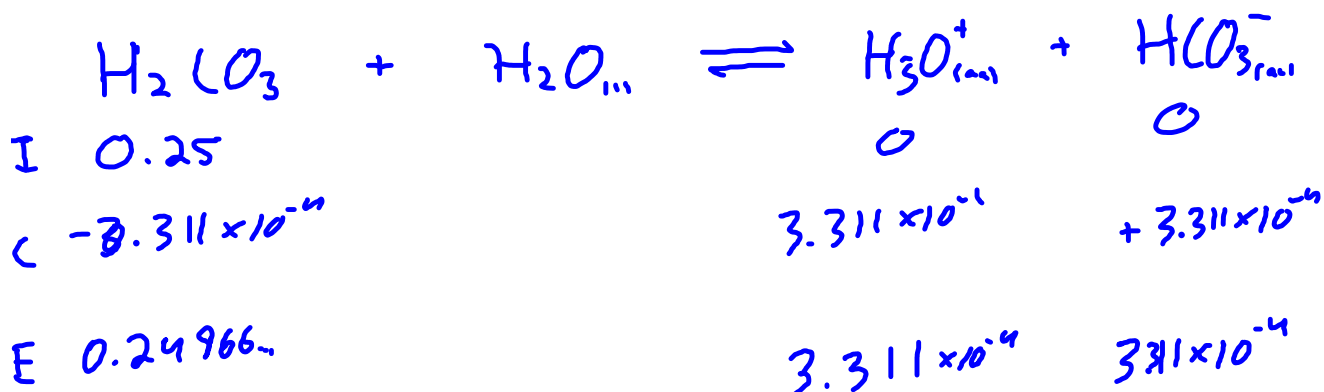
Step 4: Use your equilibrium concentrations to calculate K_a

$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]} = \frac{(0.004168)(0.004168)}{0.9958}$$

$$K_a = 1.7 \times 10^{-5}$$

Example 2:

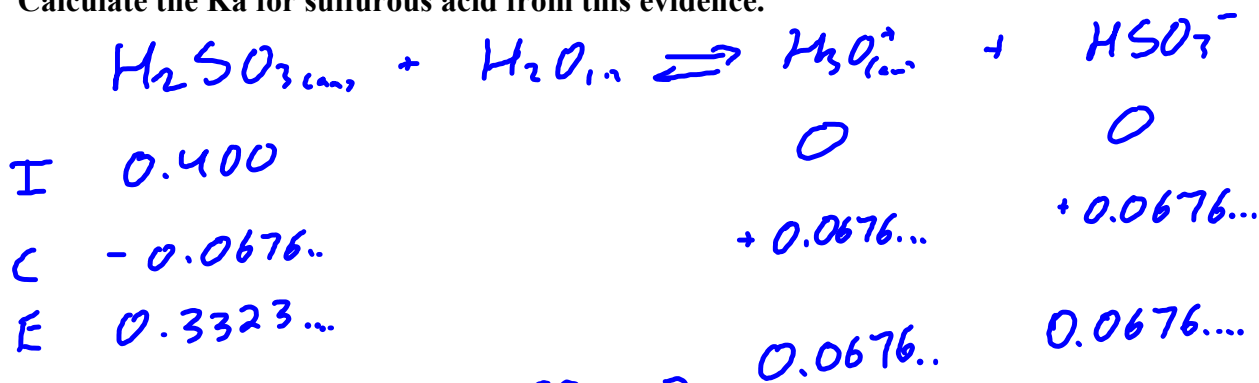
A student measures the pH of a 0.25 mol/L solution of carbonic acid to be 3.48.
Calculate the K_a for carbonic acid from this evidence.



$$K_a = \frac{[\text{HCO}_3^-][\text{H}_3\text{O}^+]}{[\text{H}_2\text{CO}_3]} = \boxed{4.4 \times 10^{-7}}$$

Example 3:

The pH of a 0.400 mol/L solution of sulfurous acid is measured to be 1.17.
Calculate the K_a for sulfurous acid from this evidence.



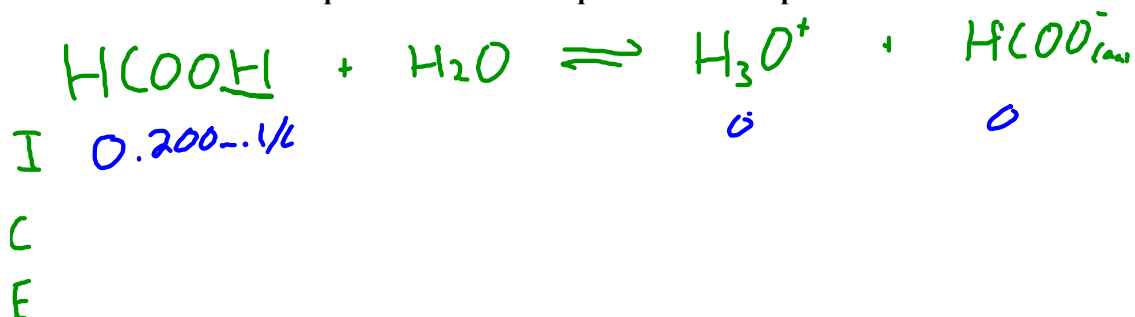
$$K_a = \frac{[\text{H}_3\text{O}^+][\text{HSO}_3^-]}{[\text{H}_2\text{SO}_3]} = \boxed{0.014}$$

Calculating $[H_3O^+]_{(aq)}$ from K_a

Example 1:

Predict the $[H_3O^+]_{(aq)}$ and pH for a 0.200 mol/L aqueous solution of methanoic acid

Use the balanced equation to write the equilibrium law expression.

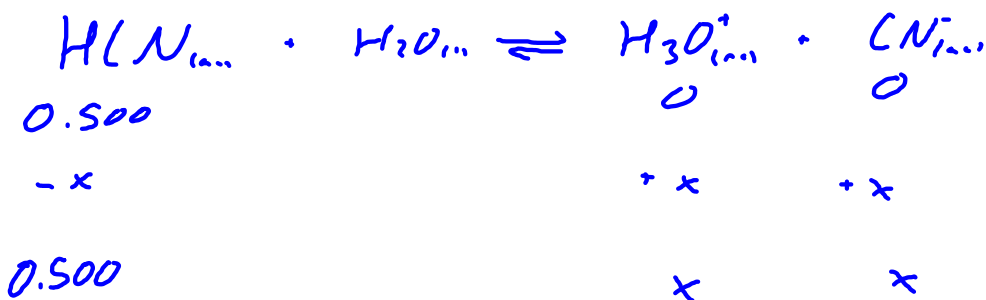


$$1.8 \times 10^{-4} = \frac{x^2}{0.200} \quad x = \sqrt{(1.8 \times 10^{-4})(0.200)}$$

$[H_3O^+] : x = 0.006 \dots /L$
 $pH = -\log(x) : 2.222$

Example 2:

Predict the $[H_3O^+]_{(aq)}$ and pH for a 0.500 mol/L aqueous solution of hydrocyanic acid.



$$K_a = \frac{[H_3O^+][CN^-]}{[HCN]} \quad 6.2 \times 10^{-10} = \frac{x^2}{0.500}$$

$[H_3O^+] : x = 1.76 \times 10^{-5} \dots /L$
 $pH = 4.754$