Using the ICE Chart to solve the PH for a weak acid or base.
Quick review:

- Weak Acids and Bases do not completely dissociate.
- We know that the equilibrium constant (Ka) can be solved for, but how can we use it to solve for the pH of a solution?
- We will need the help of our good friend Mr. ICE chart.
Ice Chart Template:

<table>
<thead>
<tr>
<th></th>
<th>Acid</th>
<th>$\text{H}^+\text{ Ion}$</th>
<th>Conjugate Base</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Initial</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>Change</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>Equilibrium</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
How to Solve for pH

- To solve for the pH of a solution from a weak acid, we will first need the $K_A$ for the particular acid.
- Each acid has its own particular $K_A$. It will be part of the question or come from a table.

*Pass out Tables Now 😊*
Getting set up:

- What is the pH of a 0.100 M solution of carbonic acid? \( K_a = 4.4 \times 10^{-7} \)
- Start with an equation and write the equilibrium expression. (\( K_a = ? \))

\[
H_2CO_3 \rightleftharpoons H_3O^+ + HCO_3^- \]

\[
K_a = \frac{[H_3O^+][HCO_3^-]}{[H_2CO_3]} \]
**Initial Concentrations**

- What is the pH of a **0.100 M solution of carbonic acid**? $K_a = 4.4 \times 10^{-7}$

- The initial concentration of the acid is given, and we can infer that no products have been created yet. They will start with zero!
### Ice Chart Template:

<table>
<thead>
<tr>
<th></th>
<th>( \text{H}_2\text{CO}_3 )</th>
<th>( \text{H}_3\text{O}^+ )</th>
<th>( \text{HCO}_3^- )</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Initial</strong></td>
<td>0.100 M</td>
<td>0.0 M</td>
<td>0.0 M</td>
</tr>
<tr>
<td><strong>Change</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>Equilibrium</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Change in Concentrations

- The **acid** will dissociate into **one** hydronium ion and **one** conjugate base.

- The acid concentration will decrease by $X$ amount. $(-X)$

- The hydronium ion and conjugate base concentrations will increase by the same $X$ amount. $(+X)$
Ice Chart Template:

<table>
<thead>
<tr>
<th></th>
<th>$\text{H}_2\text{CO}_3$</th>
<th>$\text{H}_3\text{O}^+$</th>
<th>$\text{HCO}_3^-$</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Initial</strong></td>
<td>0.100 M</td>
<td>0.0 M</td>
<td>0.0 M</td>
</tr>
<tr>
<td><strong>Change</strong></td>
<td>-X</td>
<td>X</td>
<td>X</td>
</tr>
<tr>
<td><strong>Equilibrium</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Equilibrium Concentrations

- The summation of the Initial and Change lines. Add the concentrations from the “above” boxes and write the information in the equilibrium boxes.
## Ice Chart Template:

<table>
<thead>
<tr>
<th></th>
<th>$\text{H}_2\text{CO}_3$</th>
<th>$\text{H}_3\text{O}^+$</th>
<th>$\text{HCO}_3^-$</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Initial</strong></td>
<td>0.100 M</td>
<td>0.0 M</td>
<td>0.0 M</td>
</tr>
<tr>
<td><strong>Change</strong></td>
<td>$-X$</td>
<td>$X$</td>
<td>$X$</td>
</tr>
<tr>
<td><strong>Equilibrium</strong></td>
<td>0.100M - $X$</td>
<td>$X$</td>
<td>$X$</td>
</tr>
</tbody>
</table>
What is different about this one?

- What is the pH of a 0.100 M solution of carbonic acid? $K_a = 4.4 \times 10^{-7}$
- This time we are solving for pH. To do this, we need to solve for the $[H_3O^+]$ concentration and then solve for pH.
- Use the ICE chart to solve for the $[H_3O^+]$. Use the equilibrium expression from the beginning.
Solving for \([H^+]\)

\[ \text{Ka} = \frac{[H_3O^+][HCO_3^-]}{[H_2CO_3]} \]

\[ 4.4 \times 10^{-7} = \frac{[X][X]}{[0.100 - X]} \]

So...how do we get \(X\) by itself?
Solving for \([H^+]\)

\[4.4 \times 10^{-7} = \frac{[X]^2}{[0.100 - X]}\]

*Multiply to get rid of the denominator:

\[4.4 \times 10^{-7} \cdot (0.100 - X) = X^2\]

*Distribute through the parentheses:
Solving for $[H^+]$

$4.4 \times 10^{-8} - 4.4 \times 10^{-7}X = X^2$

Look familiar yet?

*Get everything on the same side: Signs Flip!

$X^2 + 4.4 \times 10^{-7}X - 4.4 \times 10^{-8} = 0$

That’s right!!!! It’s our good friend the quadratic equation! We will need his finish solving for $X$, the $[H^+]$!!!!
Remember Quadratic?

Quadratic = \(-\frac{b + \sqrt{b^2 - 4ac}}{2a}\)

\((X^2) + (4.4 \times 10^{-7}X) + (-4.4 \times 10^{-8})\)

A \quad B \quad C

* Just plug in the numbers, not the variables. (No X’s) And Careful here! C is a negative number!!!
Plug it Into the Quadratic

\[
X = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-4.4 \times 10^{-7} \pm \sqrt{(4.4 \times 10^{-7})^2 - 4(1)(-4.4 \times 10^{-8})}}{2(1)}
\]

\[
(1)X^2 + (4.4 \times 10^{-7})X + (-4.4 \times 10^{-8}) = 0
\]

A: \(-4.4 \times 10^{-7}\)  
B: \(4.4 \times 10^{-7}\)  
C: \(-4.4 \times 10^{-8}\)

\[
X = \frac{-4.4 \times 10^{-7} \pm \sqrt{1.936 \times 10^{-13} + 1.76 \times 10^{-7}}}{2}
\]
Plug it into the Quadratic

\[ X = \frac{-4.4 \times 10^{-7} + 4.1952 \times 10^{-4}}{2} \]

\[ X = [\text{H}_3\text{O}^+] = 2.1 \times 10^{-4} \]

So now we can solve for the pH!

\[ \text{pH} = -\log (2.1 \times 10^{-4}) \]

\[ \text{pH} = 3.68 \quad \text{Woot Grats!!!} \]

You solved it! You're a genius!!!