Calculating Acid Dissociation Constants – The ICE Box Method

Problem: Calculate $K_a$ from the initial concentration of acid and the measured hydrogen ion concentration.

Example: 0.1000 mole of hydrofluoric acid (HF) is added to water to make 1 L. The hydrogen ion concentration is measured to be $7.8 \times 10^{-3}$ M (0.0078 M). What is $K_a$?

Step 1: Write balanced equation for the dissociation of the acid.

$$HF \rightleftharpoons H^+ + F^-$$

Step 2: Draw an ICE box and fill in the initial and change concentrations. For the acid, subtract Change from Initial to obtain the Equilibrium concentration:

$$0.1000 - 0.0078 = 0.0922$$

Initial – Change = Equilibrium

<table>
<thead>
<tr>
<th>Concentrations (M)</th>
<th>[HF]</th>
<th>[H$^+$]</th>
<th>[F$^-$]</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.1000</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-0.0078</td>
<td>7.8 x 10$^{-3}$</td>
<td>7.8 x 10$^{-3}$</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.0922</td>
<td>7.8 x 10$^{-3}$</td>
<td>7.8 x 10$^{-3}$</td>
</tr>
</tbody>
</table>

Step 3: Write the equilibrium constant expression and plug concentrations into the expression.

$$K_a = \frac{[H^+][F^-]}{[HF]} = \frac{7.8 \times 10^{-3} \times 7.8 \times 10^{-3}}{9.22 \times 10^{-2}} = 6.6 \times 10^{-4}$$

To calculate the percentage of acid that is dissociated into ions, divide the anion concentration by the initial acid concentration and multiply by 100:

$$\% \text{ dissociated} = \frac{[F^-]}{[HF]_{init}} \times 100 = \frac{7.8 \times 10^{-3}}{0.1000 \times 100} = 7.8\%$$

When you have calculated $K_a$ for each of the acids listed in this worksheet, write their names, ranking the acids in order of strength from the strongest (1) to the weakest (6).

1. ___________________________ $K_a =$ _______________________
2. ___________________________ $K_a =$ _______________________
3. ___________________________ $K_a =$ _______________________
4. ___________________________ $K_a =$ _______________________
5. ___________________________ $K_a =$ _______________________
6. ___________________________ $K_a =$ _______________________
A. The common name for acetylsalicylic acid is aspirin, which you can abbreviate to HA.
1. Write the chemical equation for the dissociation of aspirin.

2. Use an ICE box to calculate equilibrium concentrations if the initial concentration of aspirin is 0.1000 M and the measured hydrogen ion concentration is $5.7 \times 10^{-3}$ M.

3. What is $K_a$ for aspirin? Write the equilibrium constant expression and plug concentrations into the expression.

4. At 0.1000 M, what percentage of aspirin is dissociated into ions?

5. Is aspirin a strong acid or a weak acid?

6. What is the pH of a 0.1000 M solution of aspirin?

B. Acetic acid is the main ingredient in vinegar.
1. The chemical equation for the dissociation of acetic acid is

\[
\text{CH}_3\text{CO}_2\text{H} \rightleftharpoons \text{H}^+ + \text{CH}_3\text{CO}_2^- 
\]

2. Use an ICE box to calculate equilibrium concentrations if the initial concentration of acetic acid is 0.1000 M and the measured hydrogen ion concentration is $1.34 \times 10^{-3}$ M.

3. What is $K_a$ for acetic acid?

4. At 0.1000 M, what percentage of acetic acid is dissociated into ions?

5. Is acetic acid a strong acid or a weak acid?

6. What is the pH of a 0.1000 M solution of acetic acid?

7. What is the hydroxide ion concentration of a 0.100 M solution of aspirin?
C. Calculate $K_a$ for formic acid (HCO$_2$H) if the initial concentration of acid is 0.1000 M and the measured hydrogen ion concentration is $4.1 \times 10^{-3}$ M.

1. Write the chemical equation for the dissociation of formic acid.

2. Use an ICE box to calculate equilibrium concentrations.

3. What is $K_a$ for formic acid?

4. At 0.1000 M, what percentage of formic acid is dissociated into ions?  
   6. What is the pH of a 0.1000 M solution of formic acid?

5. Is formic acid a strong acid or a weak acid?

7. What is the hydroxide ion concentration of a 0.1000 M solution of formic acid?

D. Calculate $K_a$ for sulfamic acid (HSO$_3$NH$_2$) if the initial concentration of acid is 0.100 M and the measured hydrogen ion concentration is 0.062 M.

1. Write the chemical equation for the dissociation of sulfamic acid.

2. Use an ICE box to calculate equilibrium concentrations.

3. What is $K_a$ for sulfamic acid?

4. At 0.100 M, what percentage of sulfamic acid is dissociated into ions?  
   6. What is the pH of a 0.100 M solution of sulfamic acid?

5. Is sulfamic acid a strong acid or a weak acid?

7. What is the hydroxide ion concentration of a 0.100 M solution of sulfamic acid?
E. Calculate $K_a$ for nitric acid ($\text{HNO}_3$) if the initial concentration of acid is 1.00 M and the measured hydrogen ion concentration is 0.96 M.

1. Write the chemical equation for the dissociation of nitric acid.

2. Use an ICE box to calculate equilibrium concentrations.

3. What is $K_a$ for nitric acid?

4. At 1.00 M, what percentage of nitric acid is dissociated into ions?

6. What is the pH of a 1.00 M solution of nitric acid?

5. Is nitric acid a strong acid or a weak acid?

7. What is the hydroxide ion concentration of a 1.00 M solution of nitric acid?

F. If the initial concentration of iodic acid ($\text{HIO}_3$) is 0.100 M and the measured hydrogen ion concentration is $7.1 \times 10^{-2}$ M, what is $K_a$ for iodic acid?

1. Write the chemical equation for the dissociation of iodic acid.

2. Use an ICE box to calculate equilibrium concentrations.

3. What is $K_a$ for iodic acid?

4. At 0.100 M, what percentage of iodic acid is dissociated into ions?

6. What is the pH of a 0.100 M solution of iodic acid?

5. Is iodic acid a strong acid or a weak acid?

7. What is the hydroxide ion concentration of a 0.100 M solution of iodic acid?