Ch. 16 | Equilibria
involving acids & bases

What is an acid?
Arrhenius: when dissolved in water acids produce $H^+$ and bases $OH^-$

Braunsted-Lowry: acids donate $H^+$, bases accept protons

$\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$

*base* *acid* *acid* *base*

Conjugate pair

Lewis: acid accepts electron pairs, base donates electron pairs
pure water

\[ H_2O + H_2O \leftrightharpoons H_3O^+ + OH^- \]

(happens to a very small extent)

"auto-ionization" or "self-ionization"

water acts as both acid & base

"amphiprotic"

\[ K_w = \frac{[H_3O^+][OH^-]}{[H_2O]^2} \]

\[ K_w = 1.0 \times 10^{-14} \]

\[ \Rightarrow [H_3O^+] = [OH^-] = 1.0 \times 10^{-7} M \]

holds also in solutions.
\[ \text{pH} = -\log [\text{H}_3\text{O}^+] \leq \]

\[ \text{e.g. } [\text{H}_3\text{O}^+] = 10^{-7} \text{ m} \]

\[ \text{pH} = 7 \quad -\text{pH} \]

\[ [\text{H}_3\text{O}^+] = 10 \]

\[ \text{pOH} = -\log [\text{OH}^-] \]

\[ \text{pH} + \text{pOH} = 14 \]

Say we add acid to pure water.

Acid will dissociate \( \text{H}_3\text{O}^+ \)

\[ [\text{H}_3\text{O}^+] \uparrow \quad \text{pH} < 7 \]

\[ [\text{OH}^-] \downarrow \quad \text{pOH} > 7 \]

\[ "\text{acidic}" \]
Solutions of strong acids
dissociate completely

a) Solution of 0.25M HCl
What is the pH?

\[ HCl + H_2O \rightarrow [H_3O^+] + Cl^- \]

0.25M

\[ \text{much higher than } 10^{-7} \] (from water) negligible

so \[ [H_3O^+] = 0.25 \text{ M} \]

\[ \text{pH} = -\log (0.25) = +0.60 \]

\[ \text{pOH} = 14 - 0.60 = 13.40 \]

\[ [\text{OH}^-] = 10^{-13.40} \]
b) 0.50 M NaOH. What is the pH?

**NaOH → Na⁺ + OH⁻**

**0.50**

\[
[\text{OH}^-] = 0.50
\]

\[
\text{pOH} = -\log (0.50) = 0.30
\]

\[
\text{pH} = 14 - \text{pOH} = 13.70
\]

---

For weak acids, bases → ionization equilibrium
Weak acid:

\[
\text{CH}_3\text{COOH}^{\text{(aq)}} + \text{H}_2\text{O}^{\text{(l)}} \rightleftharpoons \text{CH}_3\text{COO}^{-}\text{(aq)} + \text{H}_3\text{O}^{+}\text{(aq)}
\]

\[
K_a = \frac{[\text{CH}_3\text{COO}^{-}] [\text{H}_3\text{O}^{+}]}{[\text{CH}_3\text{COOH}]}
\]

/ acid ionization constant

Weak base:

\[
\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^{+} + \text{OH}^{-}
\]

\[
K_b = \frac{[\text{NH}_4^{+}] [\text{OH}^{-}]}{[\text{NH}_3]}
\]

/ base ionization constant
Example

Find pH of vinegar (1 M

$K_a = 1.8 \times 10^{-5}$ (H$_3$COOH)

$K_a = \frac{[H_3O^+][C(CH_3)COO^-]}{[C2H_3COOH]}$

<table>
<thead>
<tr>
<th></th>
<th>CH$_3$COOH</th>
<th>CH$_3$COO$^-$</th>
<th>H$_3$O$^+$</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>-</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>E</td>
<td>1-x</td>
<td>x</td>
<td>x</td>
</tr>
</tbody>
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$K_a = \frac{x \cdot x}{(1.00 - x)} \Rightarrow$

$\Rightarrow x^2 + K_a x - K_a = 0$

Simplifying assumption:

$1.00 - x \approx 1.00$

$x^2 = K_a \Rightarrow x = 4.2 \times 10^{-3}$ M
Rule of thumb:

\[
\frac{\text{Molarity}}{K_a} > 100
\]

then simplification is OK.

\[
X = 4.2 \cdot 10^{-3} M
\]

\[
pH = -\log [X]
\]

\[
= -\log (4.2 \cdot 10^{-3})
\]

\[
= -\log 4.2 - \log 10^{-3}
\]

\[
= -0.62 - 3 = 2.38
\]