1. The pH of a benzoic acid (C₆H₅COOH) solution is 4.25. Calculate the concentration of the benzoic acid solution. \( K_a \) (benzoic acid) = \( 6.5 \times 10^{-5} \)

C₆H₅COOH(aq) + H₂O(l) ⇌ H₃O⁺(aq) + C₆H₅COO⁻(aq)

<table>
<thead>
<tr>
<th>Initial (M)</th>
<th>Change (M)</th>
<th>After Rxn (M)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>−x</td>
<td>1 − x</td>
</tr>
<tr>
<td>0</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>0</td>
<td>+x</td>
<td>+x</td>
</tr>
</tbody>
</table>

\[
K_a = \frac{[H_3O^+][C_6H_5COO^-]}{[C_6H_5COOH]} = \frac{x^2}{1 - x}
\]

Since the problem asks you to calculate the concentration of benzoic acid (I), it must indicate the value of \( x \). \( x \) can be calculated from the pH.

\[
pH = -\log[H_3O^+] \\
-pH = \log[H_3O^+] \\
10^{-pH} = [H_3O^+] \\
[H_3O^+] = 10^{-4.25} \\
[H_3O^+] = x = 5.6 \times 10^{-5} \text{ M}
\]

\[
6.5 \times 10^{-5} = \frac{x^2}{1 - x}
\]

\[
6.5 \times 10^{-5} = \frac{(5.6 \times 10^{-5})^2}{1 - (5.6 \times 10^{-5})} \\
6.5 \times 10^{-5}(1) - (3.64 \times 10^{-9}) = 3.136 \times 10^{-9} \\
6.5 \times 10^{-5}(1) - (3.64 \times 10^{-9}) = 6.776 \times 10^{-9} \\
1 = 1.0 \times 10^{-4} \text{ M} = [C_6H_5COOH]
\]

2. Calculate the pH of a 0.25 M solution of NH₃. \( K_b \) (ammonia) = \( 1.8 \times 10^{-5} \)

NH₃(aq) + H₂O(l) ⇌ NH₄⁺(aq) + OH⁻(aq)

<table>
<thead>
<tr>
<th>Initial (M)</th>
<th>Change (M)</th>
<th>After Rxn (M)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.25</td>
<td>−x</td>
<td>0.25 − x</td>
</tr>
<tr>
<td>0</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>0</td>
<td>+x</td>
<td>+x</td>
</tr>
</tbody>
</table>

\[
K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = \frac{x^2}{0.25 - x}
\]
Because $K$ is so small, very little reactant goes to product. We make the assumption that $0.25 - x \approx 0.25$.

\[
1.8 \times 10^{-5} = \frac{x^2}{0.25}
\]

\[x^2 = 4.5 \times 10^{-6}\]

\[x = [\text{OH}^-] = 2.12 \times 10^{-3} M\]

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

\[\text{pOH} = -\log(2.12 \times 10^{-3}) = 2.67\]

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

\[
\text{pH} = 11.33
\]

3. A 0.10 $M$ weak base has a pH of 9.12. Calculate the equilibrium constant, $K_b$, for this weak base.

\[
\text{B(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{BH}^+(aq) + \text{OH}^-(aq)
\]

<table>
<thead>
<tr>
<th>Initial ($M$)</th>
<th>0.10</th>
<th>0</th>
<th>0</th>
</tr>
</thead>
<tbody>
<tr>
<td>Change ($M$)</td>
<td>$-x$</td>
<td>$+x$</td>
<td>$+x$</td>
</tr>
<tr>
<td>After Rxn ($M$)</td>
<td>$0.10 - x$</td>
<td>$x$</td>
<td>$x$</td>
</tr>
</tbody>
</table>

\[
K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]}
\]

\[
K_b = \frac{x^2}{0.10 - x}
\]

Since the problem asks you to calculate $K_b$, it must indicate the value of $x$. $x$ can be calculated from the pH.

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

\[\text{pOH} = 14 - 9.12 = 4.88\]

\[\text{[OH}^-]\] = $10^{-\text{pOH}} = 10^{-4.88}$

\[\text{[OH}^-]\] = $1.32 \times 10^{-5} M = x$

Substitute the value of $x$ into the equilibrium constant expression to solve for $K_b$.

\[
K_b = \frac{x^2}{0.10 - x}
\]

\[
K_b = \frac{(1.32 \times 10^{-5})^2}{0.10 - (1.32 \times 10^{-5})}
\]

\[
K_b = 1.7 \times 10^{-9}
\]