1. (a) As acetic acid is a weak acid, $[\text{H}_3\text{O}^+]$ must be calculated using a reaction table:

<table>
<thead>
<tr>
<th></th>
<th>CH$_3$COOH</th>
<th>H$_2$O</th>
<th>$\rightleftharpoons$</th>
<th>H$_3$O$^+$</th>
<th>CH$_3$COO$^-$</th>
</tr>
</thead>
<tbody>
<tr>
<td>initial</td>
<td>0.15</td>
<td>large</td>
<td></td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>change</td>
<td>$-x$</td>
<td>negligible</td>
<td></td>
<td>$+x$</td>
<td>$+x$</td>
</tr>
<tr>
<td>final</td>
<td>0.15 $-x$</td>
<td>large</td>
<td></td>
<td>$x$</td>
<td>$x$</td>
</tr>
</tbody>
</table>

The equilibrium constant $K_a$ is given by:

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = \frac{x^2}{0.15 - x}$$

As $K_a = 10^{-4.76}$ is very small, $0.15 - x \sim 0.15$ and hence:

$$x^2 = 0.15 \times 10^{-4.76} \quad \text{or} \quad x = 1.6 \times 10^{-3} \text{ M} = [\text{H}_3\text{O}^+]$$

Hence, the pH is given by:

$$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+] = -\log_{10}(1.6 \times 10^{-3}) = 2.8$$

(b) Nitrous acid is also a weak acid so $[\text{H}_3\text{O}^+]$ must again be calculated:

<table>
<thead>
<tr>
<th></th>
<th>HNO$_2$</th>
<th>H$_2$O</th>
<th>$\rightleftharpoons$</th>
<th>H$_3$O$^+$</th>
<th>NO$_2$</th>
</tr>
</thead>
<tbody>
<tr>
<td>initial</td>
<td>0.018</td>
<td>large</td>
<td></td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>change</td>
<td>$-x$</td>
<td>negligible</td>
<td></td>
<td>$+x$</td>
<td>$+x$</td>
</tr>
<tr>
<td>final</td>
<td>0.018 $-x$</td>
<td>large</td>
<td></td>
<td>$x$</td>
<td>$x$</td>
</tr>
</tbody>
</table>

The equilibrium constant $K_a$ is given by:

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{HNO}_2]} = \frac{x^2}{0.018 - x}$$

As $K_a = 10^{-3.14}$ is very small, $0.018 - x \sim 0.018$ and hence:

$$x^2 = 0.018 \times 10^{-3.14} \quad \text{or} \quad x = 3.6 \times 10^{-3} \text{ M} = [\text{H}_3\text{O}^+]$$

Hence, the pH is given by:

$$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+] = -\log_{10}(3.6 \times 10^{-3}) = 2.4$$
2. The % ionisation can be calculated using:

\[
\text{percent HA dissociated} = \frac{[\text{HA}]_{\text{dissociated}}}{[\text{HA}]_{\text{initial}}} \times 100
\]

For weak acids, the concentration of acid present at equilibrium must be calculated from the reaction table. As all the acids have initial concentrations of 0.10 M, this is:

<table>
<thead>
<tr>
<th></th>
<th>HA</th>
<th>H₂O</th>
<th>⇌</th>
<th>H₃O⁺</th>
<th>A⁻</th>
</tr>
</thead>
<tbody>
<tr>
<td>initial</td>
<td>0.10</td>
<td>large</td>
<td></td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>change</td>
<td>-x</td>
<td>negligible</td>
<td>+x</td>
<td>+x</td>
<td></td>
</tr>
<tr>
<td>final</td>
<td>0.10 – x</td>
<td>large</td>
<td>x</td>
<td>x</td>
<td></td>
</tr>
</tbody>
</table>

The equilibrium constant \( K_a = \frac{[\text{H}_3\text{O}^+] [\text{A}^-]}{[\text{HA}]} = \frac{x^2}{0.10 - x} \). For very small \( K_a \) values, \( 0.10 - x \sim x \) so:

\[
x^2 = 0.10 \times K_a \quad \text{and} \quad x = \sqrt{(0.10 \times K_a)} = [\text{HA}]_{\text{dissociated}}
\]

(a) For acetic acid, \( K_a = 10^{-4.76} \) so:

\[
[\text{CH₃COOH}]_{\text{dissociated}} = \sqrt{(0.10 \times 10^{-4.76})} = 1.3 \times 10^{-3} \text{ M}
\]

\[
[\text{CH₃COOH}]_{\text{initial}} = 0.10 \text{ M}
\]

percent dissociated = \( \frac{(1.3 \times 10^{-3})}{(0.10)} \times 100 = 1.3\% \)

(b) For hydrogen cyanide, \( K_a = 10^{-9.21} \) so:

\[
[\text{HCN}]_{\text{dissociated}} = \sqrt{(0.10 \times 10^{-9.21})} = 7.8 \times 10^{-6} \text{ M}
\]

\[
[\text{HCN}]_{\text{initial}} = 0.10 \text{ M}
\]

percent dissociated = \( \frac{(7.8 \times 10^{-6})}{(0.10)} \times 100 = 0.0078\% \)

(c) For hypochlorous acid, \( K_a = 10^{-7.54} \) so:

\[
[\text{HOCl}]_{\text{dissociated}} = \sqrt{(0.10 \times 10^{-7.54})} = 5.4 \times 10^{-5} \text{ M}
\]

\[
[\text{HCN}]_{\text{initial}} = 0.10 \text{ M}
\]

percent dissociated = \( \frac{(5.4 \times 10^{-5})}{(0.10)} \times 100 = 0.054\% \)

3. For any acid and conjugate base combination in water, \( K_a \) and \( K_b \) are linked:

\[
K_a \times K_b = K_w = 10^{-14}
\]
4. Ammonia is a weak base so $[\text{H}_3\text{O}^+]$ must be calculated via the calculation of $[\text{OH}^-]$ from the reaction table:

<table>
<thead>
<tr>
<th></th>
<th>NH$_3$</th>
<th>H$_2$O</th>
<th>$\rightleftharpoons$</th>
<th>NH$_4^+$</th>
<th>OH$^-$</th>
</tr>
</thead>
<tbody>
<tr>
<td>initial</td>
<td>0.10</td>
<td>large</td>
<td></td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>change</td>
<td>-x</td>
<td>negligible</td>
<td>+x</td>
<td>+x</td>
<td></td>
</tr>
<tr>
<td>final</td>
<td>0.10 – x</td>
<td>large</td>
<td></td>
<td>x</td>
<td>x</td>
</tr>
</tbody>
</table>

The equilibrium constant $K_b$ is given by:

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = \frac{x^2}{0.10 - x}$$

As $K_b = 10^{-4.76}$ is very small, $0.10 - x \approx 0.10$ and hence:

$$x^2 = 0.10 \times 10^{-4.76} \quad \text{or} \quad x = 1.3 \times 10^{-3} \text{ M} = [\text{OH}^-]$$

Hence, the pOH is given by:

$$\text{pOH} = -\log_{10}[\text{OH}^-] = -\log_{10}[(1.3 \times 10^{-3})] = 2.9$$

Finally, as $\text{pH} + \text{pOH} = 14$, $\text{pH} = 14 - 2.9 = 11.1$

5. (a) Sodium nitrate contains the anion NO$_3^-$ which is the conjugate base of the strong acid HNO$_3$. NO$_3^-$ is therefore a very weak base and the solution will be neutral, $\text{pH} = 7$.

(a) Potassium acetate contains the acetate anion, CH$_3$CO$_2^-$, the conjugate base of the weak acid CH$_3$COOH. CH$_3$CO$_2^-$ is therefore a base and the solution will be basic, $\text{pH} > 7$, due to the equilibrium:

$$\text{CH}_3\text{CO}_2^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{CH}_3\text{COOH}(aq) + \text{OH}^-(aq)$$

(b) Ammonium chloride is acidic it contains the ammonium ion, NH$_4^+$, the conjugate acid of the weak base NH$_3$. NH$_4^+$ is therefore an acid and the solution will be acidic, $\text{pH} < 7$, due to the equilibrium:

$$\text{NH}_4^+(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_3(aq) + \text{H}_3\text{O}^+(aq)$$