. (a) As acetic acid is a weak acid, \([H_3O^+]\) must be calculated using a reaction table:

<table>
<thead>
<tr>
<th></th>
<th>CH₂COOH</th>
<th>H₂O</th>
<th>H₃O⁺</th>
<th>CH₃COO⁻</th>
</tr>
</thead>
<tbody>
<tr>
<td>initial</td>
<td>0.15</td>
<td>large</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>change</td>
<td>-x</td>
<td>negligible</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>final</td>
<td>0.15 - x</td>
<td>large</td>
<td>x</td>
<td>x</td>
</tr>
</tbody>
</table>

The equilibrium constant \(K_a\) is given by:

\[
K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]} = \frac{x^2}{0.15 - x}
\]

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

\[x^2 = 0.15 \times 10^{-4.76}\]

or \(x = 1.6 \times 10^{-3}\) M = \([H_3O^+]\)

Hence, the pH is given by:

\[pH = -\log_{10}[H_3O^+] = -\log_{10}(1.6 \times 10^{-3}) = 2.79\]

(b) Nitrous acid is also a weak acid so \([H_3O^+]\) must again be calculated:

<table>
<thead>
<tr>
<th></th>
<th>HNO₂</th>
<th>H₂O</th>
<th>H₃O⁺</th>
<th>NO₂⁻</th>
</tr>
</thead>
<tbody>
<tr>
<td>initial</td>
<td>0.018</td>
<td>large</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>change</td>
<td>-x</td>
<td>negligible</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>final</td>
<td>0.018 - x</td>
<td>large</td>
<td>x</td>
<td>x</td>
</tr>
</tbody>
</table>

The equilibrium constant \(K_a\) is given by:

\[
K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = \frac{x^2}{0.018 - x}
\]

As \(K_a = 10^{-3.14}\) is very small and the solution is not too dilute,

\[0.018 - x \sim 0.018\]

Hence:

\[x^2 = 0.018 \times 10^{-3.14}\]

or \(x = 3.6 \times 10^{-3}\) M = \([H_3O^+]\)

Hence, the pH is given by:

\[pH = -\log_{10}[H_3O^+] = -\log_{10}(3.6 \times 10^{-3}) = 2.44\]
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>H₃O⁺</th>
<th>A⁻</th>
</tr>
</thead>
<tbody>
<tr>
<td>initial</td>
<td>0.10</td>
<td>large</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>change</td>
<td>-x</td>
<td>negligible</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>final</td>
<td>0.10 – x</td>
<td>large</td>
<td>x</td>
<td>x</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 and as the solution is not too dilute, \( 0.10 - x \sim x \) 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:

\[
[\text{CH}_3\text{COOH}]_{\text{dissociated}} = \sqrt{(0.10 \times 10^{-4.76})} = 1.3 \times 10^{-3} \text{ M}
\]

\[
[\text{CH}_3\text{COOH}]_{\text{initial}} = 0.10 \text{ M}
\]

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. Ammonia is a weak base so \([\text{H}_3\text{O}^+]\) must be calculated via the calculation of \([\text{OH}^-]\) from the reaction 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:

$$\text{pOH} = -\log_{10}[\text{OH}^-] = -\log_{10}[(1.3 \times 10^{-3})] = 2.88$$

Finally, as $\text{pH} + \text{pOH} = 14$, $\text{pH} = 14.00 - 2.88 = 11.12$

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

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

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

(c) 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, pH < 7, due to the equilibrium:

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

5. This solution contains an acid (acetic acid) and its conjugate base (the acetate anion) so the Henderson-Hasselbalch equation can be used:

$$\text{pH} = pK_a + \log_{10}\left(\frac{[\text{base}]}{[\text{acid}]}\right)$$

As $[\text{acetic acid}] = [\text{sodium acetate}]$, $\log_{10}\left(\frac{[\text{base}]}{[\text{acid}]}\right) = \log_{10}(1) = 0$ and so

$$\text{pH} = pK_a = -\log_{10}(K_a) = -\log_{10}(1.8 \times 10^{-5}) = 4.74$$
6. (a) Titration of a strong acid equivalence ([acid]_{initial} = [base]_{added}) occurs at pH = 7.

The example shown below has been calculated using 50 mL of a 0.1 M solution of the acid with a 0.1 M base.

(b) Titration of a weak acid with a strong base – equivalence ([acid]_{initial} = [base]_{added}) leads to formation of the conjugate base of the weak acid and so occurs at pH > 7.
(c) Titration of a strong acid with a weak base - equivalence ([acid]_{initial} = [base]_{added}) leads to formation of the conjugate acid of the weak base and so occurs at pH < 7.

```
0 2 4 6 8 10 12 14
pH

0 20 40 60 80 100
mL of 0.1 M base

strong base: equivalence at pH = 7
weak base: equivalence at pH < 7
```