6. What is the pH and the degree of dissociation in a a) 0.1 M; in a b) 0.01 M and in a c) 0.001 M acetic acid solution, respectively?

a) Acetic acid is a weak acid, so:

$$[H^+] = K_a \cdot \frac{c_{weak\ acid} - [H^+]}{[H^+]}$$

However, if the $c_{weak\ acid} = 0.1$ M, we can use the simpler form of the formula:

$[H^+]^2 = K_a \cdot c_{weak\ acid}$

$[H^+] = \sqrt{1.86 \times 10^{-5} \times 0.10}$

$[H^+] = 1.3638 \times 10^{-3}$ M

pH = $-\log[H^+] = -\log(1.3638 \times 10^{-3}) = 2.865$

In this case $[H^+]$ = [acetate ion], so $\alpha = \frac{[H^+]}{c_{weak\ acid}} = 0.0136 = 1.36\ \%$

b)Acetic acid is a weak acid, so:

$$[H^+] = K_a \cdot \frac{c_{weak\ acid} - [H^+]}{[H^+]}$$

$[H^+]^2 = K_a(c_{weak\ acid} - [H^+])$

$[H^+]^2 + (K_a \times [H^+]) - (K_a \times c_{weak\ acid}) = 0$

By solving the equation, $[H^+] = \frac{-K_a \pm \sqrt{K_a^2 + 4 \times (K_a \times c_{weak\ acid})}}{2} = 4.22077 \times 10^{-4}$ M

pH = $-\log[H^+] = -\log(4.22077 \times 10^{-4}) = 3.375$

In this case $[H^+]$ = [acetate ion], so $\alpha = \frac{[H^+]}{c_{weak\ acid}} = 0.0422 = 4.22\ \%$