THE IONIC PRODUCT FOR WATER, K_W

1. a) $K_W = [H_3O^+]\,[OH^-]$ or $K_W = [H^+]\,[OH^-]$

b) $mol^2\ dm^{-6}$ (If you don't understand why this is, look again at the Chemguide page.)

c) Using the simpler version of the K_W expression:

$$[H^+]\,[OH^-] = 1.00 \times 10^{-14}$$

For every hydrogen ion (or hydroxonium ion) formed, there is an OH^- ion formed as well, so

$$[OH^-] = [H^+]$$

and $[H^+]^2 = 1.00 \times 10^{-14}$

so $[H^+] = 1.00 \times 10^{-7}$

$$pH = -\log_{10}[H^+] = 7$$

(If you got this wrong, check that you are entering 1.00×10^{-14} on your calculator properly, and not mis-using the EXP button. The EXP button includes the whole of "x 10" - all you need to enter is the index (or exponent - that's what EXP stands for).)

d) Using the same calculation as before:

$$[H^+]\,[OH^-] = 1.00 \times 10^{-14}$$

For every hydrogen ion (or hydroxonium ion) formed, there is an OH^- ion formed as well, so

$$[OH^-] = [H^+]$$

and $[H^+]^2 = 1.471 \times 10^{-14}$

so $[H^+] = 1.213 \times 10^{-7}$

$$pH = -\log_{10}[H^+] = 6.92$$

The water is still neutral because there are still equal numbers of hydrogen (hydroxonium) and hydroxide ions. To be acidic, it would have to have more hydrogen ions than hydroxide ions, and vice versa to be alkaline.

e) If K_W increases, there must be more hydrogen and hydroxide ions at higher temperatures. So the position of equilibrium is moving to the right as temperature increases. Le Chatelier says that a higher temperature favours the reaction which absorbs heat - the endothermic one. Therefore the forward reaction (the ionisation of the water) must be endothermic. (If you have forgotten about Le Chatelier's Principle, now might be a good time to revise it.)

2. a) $pK_W = -\log_{10} K_W$.

   b) $pK_W = -\log_{10} (1.471 \times 10^{-14}) = 13.83$