Worksheet 7.2: Solutions pH calculations	

No.	Answer
1	**a** HCl is a strong, monoprotic acid. It completely ionises in water. $\therefore [H_3O^+] = [HCl] = 0.10 = 1 \text{ x } 10^{-1}$ mol L^{-1} $\therefore pH = -\log_{10}[H_3O^+] = -\log_{10}(1 \text{ x } 10^{-1}) = 1.0$ **b** HNO_3 is a strong, monoprotic acid. It completely ionises in water. $\therefore [H_3O^+] = [HNO_3] = 0.050 = 5.0 \times 10^{-2}$ mol L^{-1} $\therefore pH = -\log_{10}[H_3O^+] = -\log_{10}(5.0 \times 10^{-2}) = 1.3$
2	**a** NaOH completely dissociates in water. $[OH^-] = 0.10 = 1 \text{ x } 10^{-1}$ $[H_3O^+] \times [OH^-] = 1 \text{ x } 10^{-14}$ $\therefore 1 \text{ x } 10^{-1} \times [H_3O^+] = 1 \text{ x } 10^{-14}$ $\therefore [H_3O^+] = 1 \text{ x } 10^{-13}$ mol L^{-1} $\therefore pH = -\log_{10}[H_3O^+] = -\log_{10}(1 \text{ x } 10^{-13}) = 13$ **b** $Ba(OH)_2$ completely dissociates in water to produce 2 OH^- ions per unit of $Ba(OH)_2$. $[OH^-] = 2 \times 0.50 = 1.0 = 1 \text{ x } 10^{0}$ $[H_3O^+] \times [OH^-] = 1 \text{ x } 10^{-14}$ $\therefore 1 \text{ x } 10^{0} \times [H_3O^+] = 1 \text{ x } 10^{-14}$ $\therefore [H_3O^+] = 1 \text{ x } 10^{-14}$ mol L^{-1} $\therefore pH = -\log_{10}[H_3O^+] = -\log_{10}(1 \text{ x } 10^{-14}) = 14$
3	**a** $[H_3O^+] = 10^{-pH} = 1 \text{ x } 10^{-11}$ mol L^{-1} $[H_3O^+] \times [OH^-] = 1 \text{ x } 10^{-14}$ $\therefore 1 \text{ x } 10^{-11} \times [OH^-] = 1 \text{ x } 10^{-14}$ $\therefore [OH^-] = 1 \text{ x } 10^{-3} = 0.0010$ mol L^{-1} **b** $[H_3O^+] = 10^{-pH} = 1 \text{ x } 10^{-2.3}$ mol L^{-1} $[H_3O^+] \times [OH^-] = 1 \text{ x } 10^{-14}$ $\therefore 1 \text{ x } 10^{-2.3} \times [OH^-] = 1 \text{ x } 10^{-14}$ $\therefore [OH^-] = 1 \text{ x } 10^{-11.7} = 2.0 \times 10^{-12}$ mol L^{-1}
4	**a** H_2SO_4 is a strong, diprotic acid. It completely ionises in water. $\therefore [H_3O^+] = 2 \times [H_2SO_4] = 2 \times 0.10 = 2.0 \times 10^{-1}$ mol L^{-1} $\therefore pH = -\log_{10}[H_3O^+] = -\log_{10}(2.0 \times 10^{-1}) = 0.7$ **b** H_2SO_4 is a strong, diprotic acid, but only one hydrogen ion is completely donated in water. $\therefore [H_3O^+] = [H_2SO_4] = 0.10 = 1 \text{ x } 10^{-1}$ mol L^{-1} $\therefore pH = -\log_{10}[H_3O^+] = -\log_{10}(1 \text{ x } 10^{-1}) = 1.0$ **c** The first hydrogen ion is completely donated in water. The second hydrogen ion is only partially donated. Therefore the pH will be intermediate between 0.7 and 1.0. The exact pH will depend on the extent of ionisation of the second hydrogen ion.

5	$n(H^+) = c \times V = 1 \text{ x } 10^{-3} \times 0.0200 = 2.00 \times 10^{-5}$ mol $[H_3O^+] = \frac{n}{V} = \frac{2.00 \times 10^{-5}}{0.200} = 1 \text{ x } 10^{-4}$ mol L^{-1} pH = $-\log[H_3O^+] = -\log(1 \text{ x } 10^{-4}) = 4.0$ A 10-fold dilution produces a change of one pH unit.
6	$c_1V_1 = c_2V_2$ $c_1 = [H_3O^+] = 10^{-pH} = 1 \text{ x } 10^{-2}$ and $c_2 = [H_3O^+] = 10^{-pH} = 1 \text{ x } 10^{-2.5}$ $\therefore 1 \text{ x } 10^{-2} \times 50.0 = 1 \text{ x } 10^{-2.5} \times c_2$ $\therefore c_2 = 158$ mL, therefore 108 mL must be added
7	The solutions listed in order of decreasing pH are: NaOH: a strong base $\therefore$ high pH NH_3: a weak base $\therefore$ pH > 7 but not too high H_2O: pH = 7 CH_3COOH: a weak acid $\therefore$ pH < 7 HNO_3: a strong acid $\therefore$ low pH H_2SO_4: a strong, diprotic acid $\therefore$ pH < pH of HNO_3
8	$n(H^+)$ for first solution = $c \times V = 1 \text{ x } 10^{-5} \times 0.0250 = 2.50 \times 10^{-7}$ mol $n(H^+)$ for second solution = $c \times V = 1 \text{ x } 10^{-6} \times 0.0250 = 2.50 \times 10^{-8}$ mol $n(H^+)$ for final solution = $2.5 \times 10^{-7} + 2.5 \times 10^{-8} = 2.75 \times 10^{-7}$ mol $[H_3O^+]$ for final solution = $\frac{n}{V} = \frac{2.75 \times 10^{-7}}{0.0500} = 5.5 \times 10^{-6}$ mol L^{-1} pH = $-\log[H_3O^+] = -\log(5.5 \times 10^{-6}) = 5.3$