Worksheet 7.2 **pH calculations**	

NAME: **CLASS:**

INTRODUCTION

- The pH scale is used as a measure of the acidity or basicity of a solution. The scale is usually applied over the range 0 to 14 (but does extend beyond these values).
- $pH = -\log_{10}[H_3O^+]$
- pH is a logarithmic scale, so a difference of one unit on the pH scale means a ten-fold difference in the hydrogen ion concentration.
- For dilute solutions at 25°C, $K_W = [H_3O^+] \times [OH^-] = 1 \text{ x } 10^{-14}$
- In acidic solutions $[H_3O^+] > [OH^-]$ hence $[H_3O^+] > 1 \text{ x } 10^{-7}$ hence $pH < 7$ (at 25°C)
  In neutral solutions $[H_3O^+] = [OH^-]$ hence $[H_3O^+] = 1 \text{ x } 10^{-7}$ hence $pH = 7$ (at 25°C)
  In basic solutions $[H_3O^+] < [OH^-]$ hence $[H_3O^+] < 1 \text{ x } 10^{-7}$ hence $pH > 7$ (at 25°C)

(Assume all the calculations below are for solutions at 25°C)

No.	Question	Answer
1	Calculate the hydronium ion concentration and the pH of a: **a** 0.10 mol L^{-1} HCl solution **b** 0.050 mol L^{-1} HNO_3 solution.	
2	Calculate the hydroxide ion concentration and the pH of a: **a** 0.10 mol L^{-1} NaOH solution **b** 0.50 mol L^{-1} $Ba(OH)_2$ solution.	
3	Calculate the hydronium ion and hydroxide ion concentrations in: **a** an ammonia cleaner with a pH of 11.0 **b** lemon juice with a pH of 2.3.	

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4	Sulfuric acid is a strong diprotic acid. One student calculated the expected pH of a 0.10 mol L^{-1} H_2SO_4 solution to be 0.7, while another calculated it to be 1.0. The actual pH was found to be between these two values. **a** Show how a pH of 0.7 was calculated. **b** Show how a pH of 1.0 was calculated. **c** Explain why the actual pH was between the two calculated values.	
5	20.0 mL of a solution of pH 3.0 is diluted to produce a total volume of 200.0 mL. What is the pH of the resulting solution?	
6	What volume of water must be added to 50.0 mL of a hydrochloric acid solution of pH 2.0 to increase the pH to 2.5?	
7	List the following 1.0 mol L^{-1} solutions in order of decreasing pH. Give reasons for your order. $NaOH$, H_2O, NH_3, CH_3COOH, H_2SO_4, HNO_3	
8	25.0 mL of a solution of pH 5.0 is added to 25.0 mL of a solution of pH 6.0. What is the pH of the resultant solution?	

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} $pH = -\log[H_3O^+] = -\log(5.5 \times 10^{-6}) = 5.3$

Worksheet 7.3 pH and K_w	

NAME: **CLASS:**

INTRODUCTION

pH is used as a measure of the acidity or basicity of a solution, and is calculated using the following formula:

$$pH = -\log_{10}[H_3O^+]$$

Transposing this formula gives the formula:

$$[H_3O^+] = 10^{-pH}$$

It is found that in aqueous solutions at 25°C:

$$[H_3O^+][OH^-] = 1 \times 10^{-14}$$

Using this information we are able to calculate the pH of both acidic and basic solutions at 25°C.

No.	Question	Answer
1	Complete the table below by placing either an =, > or < symbol in the third column. (The first row has been completed as an example.) All statements refer to solutions at 25°C.	

In acidic solutions	$[H_3O^+]$	>	$[OH^-]$
In basic solutions	$[H_3O^+]$		1×10^{-7} mol L^{-1}
In neutral solutions	$[OH^-]$		1×10^{-7} mol L^{-1}
In acidic solutions	$[H_3O^+]$		1×10^{-7} mol L^{-1}
In basic solutions	$[H_3O^+]$		$[OH^-]$
In acidic solutions	$[OH^-]$		1×10^{-7} mol L^{-1}
In basic solutions	$[OH^-]$		1×10^{-7} mol L^{-1}

No.	Question	Answer
2	Explain why, in pure water, the molar concentrations of H_3O^+ and OH^- are equal.	

Worksheet 7.3

pH and K_w

3	Rain from clean air at 25°C has a pH of approximately 6, but acid rain may have a pH as low as 3. **a** With the aid of equations, explain why normal rainwater is acidic. **b** How many times more acidic is acid rain with a pH of 3 than normal rainwater with a pH of 6?	
4	**a** Calculate the pH of: **i** a 0.50 mol L^{-1} HNO_3 solution **ii** a 0.050 mol L^{-1} $Ba(OH)_2$ solution. **b** Give two reasons why two acid solutions of equal concentration could have different pH values.	

Water undergoes self-ionisation according to the equation:

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq); \qquad \Delta H = +57 \text{ kJ}$$

As this is an equilibrium reaction, the equilibrium constant for the reaction will be temperature dependent, and hence the $[H_3O^+]$ of pure water will be temperature dependent. This, in turn, means that the pH of pure water will be temperature dependent, so it is possible for a neutral solution to have a pH other than 7! The following questions consider this concept.

No.	Question	Answer
5	Write the equilibrium law expression for the self-ionisation of water.	
6	Given that the density of water is 1.0 g mL^{-1}, determine the molar concentration ($[H_2O]$) of water.	

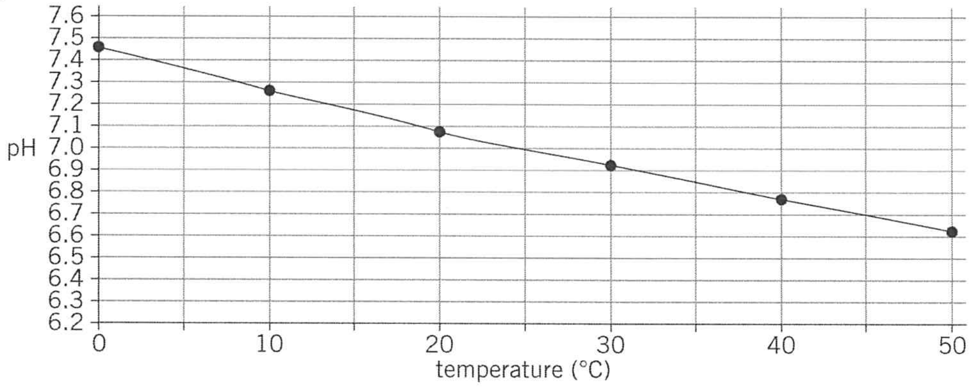
Worksheet 7.3

pH and K_w

7	Given that the $[H_2O]$ in pure water is so large, it may be considered to be constant, and therefore the equilibrium expression may be written as: $K_w = K \times [H_2O]^2 = [H_3O^+][OH^-]$ What effect will an increase in temperature have on: **a** the value of K_w? **b** the pH of pure water? **c** the neutrality of pure water? Explain your answers.	

The change in pH of water with temperature is shown in the graph below. Use this graph to answer question **8**.

<<Insert Figure 7.3>>



No.	**Question**	**Answer**
8	Determine: **a** the $[H_3O^+]$ in pure water at 45°C **b** the $[OH^-]$ in pure water at 35°C **c** K_w for pure water at 18°C.	

Worksheet 7.3: Solutions
pH and K_w

No.	Answer
1	(see table below)

In acidic solutions	$[H_3O^+]$	>	$[OH^-]$
In basic solutions	$[H_3O^+]$	<	10^{-7} mol L^{-1}
In neutral solutions	$[OH^-]$	=	10^{-7} mol L^{-1}
In acidic solutions	$[H_3O^+]$	>	10^{-7} mol L^{-1}
In basic solutions	$[H_3O^+]$	<	$[OH^-]$
In acidic solutions	$[OH^-]$	<	10^{-7} mol L^{-1}
In basic solutions	$[OH^-]$	>	10^{-7} mol L^{-1}

2 Water undergoes self-ionisation according to the following equation:

$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$

As the mole ratio of $H_3O^+(aq)$ to $OH^-(aq)$ in this equation is 1:1, the molar concentrations of $H_3O^+(aq)$ and $OH^-(aq)$ must remain equal in pure water.

3

a $CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$

$H_2CO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$

b 1000 times (The pH scale is logarithmic, and a change of 1 pH unit equals a 10-fold change in $[H_3O^+]$.)

4

a i HNO_3 is a strong, monoprotic acid. It completely ionises in water.

$\therefore [H_3O^+]=[HNO_3]=0.50 \text{ mol L}^{-1}$

$\therefore \text{pH} = -\log_{10}[H_3O^+] = -\log_{10}(0.50) = 0.3$

ii $Ba(OH)_2$ completely dissociates in water to produce 2 OH^- ions per unit of $Ba(OH)_2$.

$[OH^-]= 2 \times 0.050=0.10= 10^{-1} \text{ mol L}^{-1}$

$[H_3O^+]\times[OH^-]=10^{-14}$

$\therefore 10^{-1} \times [H_3O^+]=10^{-14}$

$\therefore [H_3O^+]=10^{-13} \text{mol L}^{-1}$

$\therefore \text{pH} = -\log_{10}[H_3O^+] = -\log_{10}10^{-13} = 13$

b They could have different acid strengths (e.g. one strong, one weak) or be different acid types (e.g. one monoprotic, one diprotic).

5 $K_w = [H_3O^+]\,[OH^-]$

6 1.0 g in 1.0 mL, $\therefore$ 1000 g in 1.0 L, $\therefore \dfrac{1000}{18.016} = 55.5$ mol in 1.0 L

$\therefore$ 56 mol L^{-1}

Worksheet 7.3: Solutions

pH and K_w

No.	Answer
7	**a** As this is an endothermic reaction in the forward direction, an increase in temperature will cause the reaction to proceed forward, thus increasing the concentrations of both H_3O^+ and OH^- ions. Therefore the equilibrium constant, K_w, will increase with an increase in temperature. **b** As pH is defined in terms of the concentration of H_3O^+ ions, an increase in temperature results in a higher concentration of H_3O^+ and so a lower pH value for pure water. **c** The concentrations of both H_3O^+ and OH^- will increase to precisely the same extent with an increase in temperature, as can be seen from the stoichiometric 1:1 ratio in the equation. Thus, although both K_w and pH change with temperature, the water will remain neutral. (Pure water is always neutral, but it only has a pH of precisely 7 at 25°C.)
8	**a** $[H_3O^+] = 10^{-pH} = 10^{-6.7} = 2.0 \times 10^{-7}$ mol L^{-1} **b** $[OH^-] = [H_3O^+] = 10^{-pH} = 10^{-6.85} = 1.4 \times 10^{-7}$ mol L^{-1} **c** $K_w = [H_3O^+][OH^-] = [H_3O^+]^2 = (10^{-pH})^2 = (10^{-7.1})^2 = 10^{-14.2} = 6.3 \text{ x } 10^{-15}$