| Worksheet 7.3 pH and $K_w$ | |
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**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. | |
| **2** | Explain why, in pure water, the molar concentrations of $H_3O^+$ and $OH^-$ are equal. | |

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

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| **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.<br>**a** With the aid of equations, explain why normal rainwater is acidic.<br>**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:<br>**i** a 0.50 mol $L^{-1}$ $HNO_3$ solution<br>**ii** a 0.050 mol $L^{-1}$ $Ba(OH)_2$ solution.<br>**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. | |

| | | |
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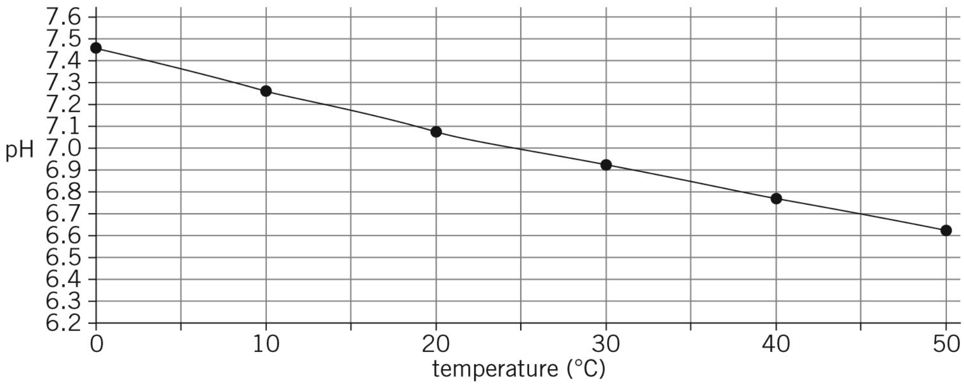
| **7** | Given that the $[H_2O]$ in pure water is so large, it may be considered to be constant, and therefore the equilibrium expression may be written as:<br>$K_w = K \times [H_2O]^2 = [H_3O^+][OH^-]$<br>What effect will an increase in temperature have on:<br>**a** the value of $K_w$?<br>**b** the pH of pure water?<br>**c** the neutrality of pure water?<br>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:<br>**a** the $[H_3O^+]$ in pure water at 45°C<br>**b** the $[OH^-]$ in pure water at 35°C<br>**c** $K_w$ for pure water at 18°C. | |