Hydrogen ion concentration (A2)

Hydrogen ion concentration determines the acidity of a solution.

Hydroxide ion concentration determines the alkalinity of a solution.

For strong acids and strong bases, the concentration of ions is very much larger than their weaker counterparts which only partially dissociate.

pH

definition: the negative logarithm to the base 10 of the hydrogen ion concentration in $mol\,dm^{-3}$.

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

to convert pH into hydrogen ion concentration,

$$[H^+] = \text{antilog}_{10}(-pH)$$

pOH

an equivalent calculation for bases converts the hydroxide ion concentration to pOH.

$$pOH = -\log_{10}[OH^-]$$

Note : [ ] represents the concentration in $mol\,dm^{-3}$

Calculating pH values from $[H^+]$

Example:

Calculate the pH of a solution whose H^+ ion concentration is 5.32×10^{-4} mol dm^{-3}.

Working,

$$pH = -\log_{10}[H^+]$$
$$= -\log_{10}(5.32 \times 10^{-4})$$
$$= 3.27$$

Exercise

Calculate the pH of the following solutions:

a. $[H^+] = 3.00 \times 10^{-4}$ mol dm^{-3}

b. $[H^+] = 1.00 \times 10^{-2}$ mol dm^{-3}

c. $[H^+] = 4.00 \times 10^{-8}$ mol dm^{-3}

d. $[H^+] = 5.40 \times 10^{-12}$ mol dm^{-3}

e. $[H^+] = 7.80 \times 10^{-10}$ mol dm^{-3}

Answers

a. 3.5　b. 2.0　c. 7.4　d. 11.3　e. 9.1

Calculating $[H^+]$ from pH

Example.

Calculate the hydrogen ion concentration of a solution whose pH is 10.5.

Working.

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

$$[H^+] = \text{antilog}_{10}(-pH)$$

$$= \text{antilog}_{10}(-10.5)$$

$$= 3.16 \times 10^{-11} \text{ mol dm}^{-3}$$

Exercise

Calculate the concentration of hydrogen ions in solutions having the following pH values:

a. pH 2.90

b. pH 3.70

c. pH 11.20

d. pH 5.40

e. pH 12.90

Answers

a. 1.26×10^{-3} mol dm^{-3}

b. 2.00×10^{-4} mol dm^{-3}

c. 6.31×10^{-12} mol dm^{-3}

d. 3.98×10^{-6} mol dm^{-3}

e. 1.26×10^{-13} mol dm^{-3}

The relationship between pH and pOH

Because H^+ and OH^- ions are produced in equal amounts when water dissociates, their concentrations will be the same.

$$[H^+] = [OH^-] = 10^{-7}\ mol\,dm^{-3}$$

From Kw equation,

$$[H^+][OH^-] = 10^{-14}\ mol^2 dm^{-6}$$

take logs on both sides,

$$\log_{10}[H^+] + \log_{10}[OH^-] = -14$$

multiply by minus,

$$-\log_{10}[H^+] - \log_{10}[OH^-] = 14$$

change to pH and pOH,

$$\boxed{pH + pOH = 14}$$

As they are based on the position of equilibrium and that varies with temperature, the above values are only true if the temperature is 25°C.

Neutral solutions may be regarded as those where $[H^+] = [OH^-]$.

Therefore a neutral solution is pH 7 only at a temperature of 25°C.

The value of K_w is constant for any aqueous solution at the stated temperature.

$[H^+]$	1	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}	10^{-9}	10^{-10}	10^{-11}	10^{-12}	10^{-13}	10^{-14}
$[OH^-]$	10^{-14}	10^{-13}	10^{-12}	10^{-11}	10^{-10}	10^{-9}	10^{-8}	10^{-7}	10^{-6}	10^{-5}	10^{-4}	10^{-3}	10^{-2}	10^{-1}	1
pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
		strongly acidic			weakly acidic			neutral		weakly alkaline				strongly alkaline	

The pH of strong acids

Monobasic acids contain only one replaceable hydrogen atom per molecule.

Strong monobasic acids such as hydrochloric acid are completely ionised in solution.

The concentration of hydrogen ions in solution is approximately the same as the concentration of the acid.

Example:

$HCl_{(aq)}$ at various concentrations.

concentration / $mol\,dm^{-3}$	pH
0·1	$-\log_{10}(1\times10^{-1}) = 1$
0·01	$-\log_{10}(1\times10^{-2}) = 2$
0·001	$-\log_{10}(1\times10^{-3}) = 3$

Diluting the acid 10 times reduces the value of the H^+ ion concentration by one tenth and increases the pH by a value of one

Calculating the pH of strong bases

Strong bases (eg. NaOH) ionise completely in solution.

The concentration of hydroxide ions in a solution of sodium hydroxide is approximately the same as the concentration of the sodium hydroxide.

To calculate the pH of a solution of strong base, the following details are required:

- the concentration of OH^- ions in solution.
- the equilibrium expression for the ionisation of water: $K_w = [H^+][OH^-]$
- the value of K_w for water.

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

$$[H^+] = \frac{K_w}{[OH^-]}$$

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

Example:

Calculate the pH of a solution of sodium hydroxide of concentration 0.0500 $mol\,dm^{-3}$

$K_w = 1.00 \times 10^{-14}\, mol^2\,dm^{-6}$ (at 298K)

Working

Step 1 Write the expression relating $[H^+]$ to K_w and $[OH^-]$

$$[H^+] = \frac{K_w}{[OH^-]}$$

Step 2 Substitute the values into the expression to calculate $[H^+]$

$$[H^+] = \frac{1.00 \times 10^{-14}}{0.0500} = 2.00 \times 10^{-13} \text{ mol dm}^{-3}$$

Step 3 Calculate the pH

$$pH = -\log_{10}[H^+]$$
$$= -\log_{10}(2.00 \times 10^{-13})$$
$$= 12.7$$

Alternatively

$$pOH = -\log_{10}[OH^-]$$
$$= -\log_{10}(0.0500)$$
$$= 1.3$$

since $pH + pOH = 14$

$$pH = 14 - pOH$$
$$= 12.7$$

Exercise

Find the pH of the following strong acids and strong bases:

a. 1.00 mol dm^{-3} HNO_3

b. $0.500 \text{ mol dm}^{-3}$ HNO_3

c. an aqueous solution containing 3.00 g HCl per dm^3

d. $0.00100 \text{ mol dm}^{-3}$ KOH ($K_w = 1.00 \times 10^{-14} \text{ mol}^2 \text{dm}^{-6}$)

e. an aqueous solution containing 0.200 g of NaOH per dm^3 ($K_w = 1.00 \times 10^{-14} \text{ mol}^2 \text{dm}^{-6}$)

Workings.

a. $pH = -\log_{10}[H^+] = -\log_{10}(1 \times 10^0) = 0$

b. $pH = -\log_{10}[H^+] = -\log_{10}(0.500) = 0.32$

c. $M_r(HCl) = 36.5$

$n(HCl) = \frac{3.00}{36.5} = 0.0822 \text{ mol}$

$[H^+] = [HCl] = 0.0822 \text{ mol dm}^{-3}$

$pH = -\log_{10}(0.0822) = 1.09$

d. $KOH_{(aq)} \longrightarrow K^+_{(aq)} + OH^-_{(aq)}$

$[OH^-] = 1 \times 10^{-3} \text{ mol dm}^{-3}$

$K_w = [H^+][OH^-] = 1 \times 10^{-14} \text{ mol}^2 \text{dm}^{-6}$

$[H^+] = \frac{1 \times 10^{-14}}{[OH^-]} = \frac{1 \times 10^{-14}}{1 \times 10^{-3}} = 1 \times 10^{-11} \text{ mol dm}^{-3}$

$pH = -\log_{10}[H^+] = -\log_{10}(1 \times 10^{-11}) = 11$

e. $M_r(NaOH) = 23 + 16 + 1 = 40$

$$n(NaOH) = \frac{0.200}{40} = 5.00 \times 10^{-3}\ mol$$

$$NaOH_{(aq)} \rightleftharpoons Na^+_{(aq)} + OH^-_{(aq)}$$

$$[OH^-] = [NaOH] = 5.00 \times 10^{-3}\ mol\,dm^{-3}$$

$$K_w = [H^+][OH^-] = 1.00 \times 10^{-14}\ mol^2 dm^{-6}$$

$$[H^+] = \frac{1.00 \times 10^{-14}}{[OH^-]} = \frac{1.00 \times 10^{-14}}{5.00 \times 10^{-3}} = 2.00 \times 10^{-12}\ mol\,dm^{-3}$$

$$pH = -\log_{10}[H^+] = -\log_{10}(2.00 \times 10^{-12}) = 11.7$$