Get some extra practice…

…working with pH and buffer solutions

1.   What is the pH of a solution containing 3.5×10^{-3} M HCl?

2.   What is the pH of a 4.74×10^{-5} mmol L^{-1} sodium hydroxide solution?

3.   What is the hydronium ion concentration of urine at pH 5.2?

4.   What is the hydroxide ion concentration in a solution if the hydronium ion concentration is 5.67×10^{-3} mol L^{-1}?

5.   Calculate the hydronium ion concentration of a solution with a pH of -0.26.

6.   What is the pH of a solution concentration, in mol L^{-1}, of 4.19 mL of a solution containing 0.156 g of hydrogen peroxide?

7.   Ethanoic acid has a K_a of 1.8×10^{-5}. What is the pH of a solution containing 0.24 M ethanoic acid and 0.13 M of the ethanoate anion?

8.   What is the ratio of lactic acid to lactate in a lactate buffer solution at pH 4.8 if the K_a of lactic acid is 1.38×10^{-4}?

9.   A weak acid has a pK_a of 7.89. Calculate the pH of a solution in which the ratio of the concentration of dissociated acid to acid anion (HA/A^-) equals 2.92.

10.   A weak acid has a pK_a of 7.07. If the pH of the solution is 7.38, what percentage of the acid is undissociated?

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11.   **Answers**

1.   What is the pH of a solution containing 3.5×10^{-3} M HCl?

In order to determine the pH of a solution, you need to know the concentration of hydrogen ions in the solution. You are not given this in the question, but you are given the concentration of the acid, HCl. Hydrochloric acid is a strong acid, so we we know that it is fully ionised in water. In other words, 3.5×10^{-3} M HCl yields 3.5×10^{-3} M hydrogen ions upon dissociation.

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

So $pH = -\log_{10}(3.5 \times 10^{-3}) = $ **2.46.**

2.   What is the pH of a 4.74×10^{-5} mmol L^{-1} sodium hydroxide solution?

Sodium hydroxide is completely dissociated in water such that 4.74×10^{-5} mmol sodium hydroxide yields 4.74×10^{-5} mol OH$^-$ ions per litre.

The ion product of water states that

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

We can rearrange this equation to isolate our unknown term, $[H_3O^+]$:

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

We know that $[OH^-] = 4.74 \times 10^{-5}$ mM $= (4.74 \times 10^{-5}) \times 10^{-3}$ M $= 4.74 \times 10^{-8}$ M.

So we can write that

$$\left[H_3O^+ \right] = \frac{1 \times 10^{-14}}{4.74 \times 10^{-8}} = 2.11 \times 10^{-7} M$$

We then use the relationship $pH = -\log_{10} [H_3O^+]$

$$pH = -\log_{10}\left(2.11 \times 10^{-7} \right)$$
$$= 6.68$$

3.   What is the hydronium ion concentration of urine at pH 5.2?

We start with the relationship $pH = -\log_{10} [H_3O^+]$. In this case, the pH is 5.2, so we can write $5.2 = -\log_{10} [H_3O^+]$.

One of the rules of logarithms is that, if $y = \log_{10} x$, $x = 10^y$.

In this case, if $5.2 = -\log_{10} [H_3O^+]$, then $-5.2 = \log_{10} [H_3O^+]$. We can then use the rule above to say that $[H_3O^+] = 10^{-5.2} = 6.3 \times 10^{-6}$ mol L^{-1}.

4.   What is the hydroxide ion concentration in a solution if the hydronium ion concentration is 5.67×10^{-3} mol L^{-1}?

Using the relationship $K_w = [H_3O^+][OH^-]$ and rearranging to solve for $[OH^-]$:

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1 \times 10^{-14}}{5.67 \times 10^{-3}}$$

$$= 1.76 \times 10^{-11} \text{ mol } L^{-1}$$

5.   Calculate the hydronium ion concentration of a solution with a pH of -0.26.

A negative pH suggests a hydronium concentration of greater than 1 M.

$pH = -0.26 = -\log[H_3O]^+$

Taking antilogs of both sides: $[H_3O^+] = 10^{-(-0.42)} = 1.8$ M

6.   What is the concentration, in mol L^{-1}, of 4.19 mL of a solution containing 0.156 g of hydrogen peroxide, H_2O_2?

The molar mass of H_2O_2 is 34 g mol^{-1} – that is, there is 1 mol of H_2O_2 in 34 g.

The amount of substance present, in mol $= \dfrac{\text{mass of sample}}{\text{molar mass}}$

$\dfrac{0.156 \text{ g}}{34 \text{ g mol}^{-1}} = 0.00459$ mol

The concentration of a solution $= \dfrac{\text{amount of substance (mol)}}{\text{volume (L)}}$

In this case, the volume of solution = 0.00419 L (4.19 mL = 0.00419 L).

So concentration of solution $= \dfrac{0.00459 \text{ mol}}{0.00419 \text{ L}} = 1.095$ M

7.   Ethanoic acid has a K_a of 1.8×10^{-5}. What is the pH of a solution containing 0.24 M ethanoic acid and 0.13 M of the ethanoate anion?

We can use the Henderson-Hasselbalch equation to calculate the pH:

$$pH = pK_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

First, we calculate the value of pK_a from the K_a:

$pK_a = -\log_{10} K_a$

$= -\log_{10} (1.8 \times 10^{-5}) = 4.74$

We can then return to the Henderson-Hasselbalch equation:

$$pH = pK_a + \log \frac{[\text{base}]}{[\text{acid}]}$$
$$= 4.74 + \log \frac{0.13}{0.24}$$
$$= 4.74 + (-0.266)$$
$$= 4.47$$

8.   What is the ratio of lactic acid to lactate in a lactate buffer solution at pH 4.8 if the K_a of lactic acid is 8.3×10^{-4}?

This is another situation in which we can use the Henderson-Hasselbalch equation:

$$pH = pK_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

First, we calculate the value of pK_a from the K_a:

$pK_a = -\log_{10} K_a$

$= -\log_{10} 8.3 \times 10^{-4} = 3.08$

We then feed the values of pH and pK_a into the Henderson-Hasselbalch equation:

$$pH = pK_a + \log\frac{[\text{base}]}{[\text{acid}]}$$

$$4.8 = 3.08 + \log\frac{[\text{base}]}{[\text{acid}]}$$

$$4.8 - 3.08 = \log\frac{[\text{base}]}{[\text{acid}]}$$

$$1.72 = \log\frac{[\text{base}]}{[\text{acid}]}$$

We then take antilogs of both sides:

$$10^{1.72} = \frac{[\text{base}]}{[\text{acid}]} = 52.48$$

So the ratio of lactate:lactic acid is 52.48:1.

9.   A weak acid has a pK_a of 7.89. Calculate the pH of a solution in which the ratio of the concentrations of dissociated acid to acid anion (HA/A⁻) equals 2.92.

We can use the Henderson-Hasselbalch equation and the properties of logarithms.

$$\text{If } pH = pK_a + \log\frac{[\text{base}]}{[\text{acid}]}$$

$$\text{then } pH = pK_a - \log\frac{[\text{acid}]}{[\text{base}]}$$

$$pH = 7.89 - \log 2.92$$

$$= 7.89 - 0.465$$

$$= 7.42$$

10.   A weak acid has a pK_a of 7.07. If the pH of the solution is 7.38, what percentage of the acid is undissociated?

To answer this question, we need to use the acid dissociation equation that relates to this equilibrium reaction:

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

$$K_a = \frac{[A^-][H_3O^+]}{[HA]}$$

Remember, because it is a large value and constant, we can leave $[H_2O]$ out of the equation.

We now need to calculate a value for K_a. $pK_a = -\log_{10} K_a$. So, if pK_a is 7.07 then, using the rules of logarithms, $K_a = 10^{-7.07} = 8.51 \times 10^{-8}$

We also know that the pH is 7.38. pH is related to the concentration of hydronium ions, $[H_3O^+]$, by the relationship $pH = -\log [H_3O^+]$. Using the rules of logarithms:

$pH = -\log [H_3O^+]$

$7.38 = -\log [H_3O^+]$

$10^{-7.38} = [H_3O^+] = 4.17 \times 10^{-8}$

The stoichiometry of the reaction, $HA + H_2O \rightleftharpoons H_3O^+ + A^-$, tells us that H_3O^+ and A^- form in equal amounts as a result of the dissociation. That is, if $[H_3O^+]$ is 4.17×10^{-8}, $[A^-]$ must take the same value.

So, if we return to the original acid dissociation equation

$$K_a = \frac{[A^-][H_3O^+]}{[HA]}$$

we can now say that

$$8.51 \times 10^{-8} = \frac{(4.17 \times 10^{-8})(4.17 \times 10^{-8})}{[HA]}$$

We can now rearrange the equation to solve for [HA].

$$[HA] = \frac{(4.17 \times 10^{-8})(4.17 \times 10^{-8})}{(8.51 \times 10^{-8})}$$

$$= 2.04 \times 10^{-8} M$$

If 4.17×10^{-8} mol of the acid has dissociated, and 2.04×10^{-8} mol remains undissociated, the total amount of acid present must be $(4.17 \times 10^{-8}) + (2.04 \times 10^{-8}) = 6.21 \times 10^{-8}$ mol.

Therefore the percentage of undissociated acid = $(2.04 \times 10^{-8})/(6.21 \times 10^{-8})\% =$

32.9%

Alternatively you can use the Henderson-Hasselbalch equation:

$$pH = pK_a + \log\frac{[A^-]}{[HA]}$$

$$7.38 = 7.07 + \log\frac{[A^-]}{[HA]}$$

$$\log\frac{[A^-]}{[HA]} = 7.38 - 7.07 = 0.31$$

Taking inverse logs of both sides:

$$\frac{[A^-]}{[HA]} = 10^{0.31} = 2.04$$

If we allow the percentage of undissociated acid HA to be $x\%$, then the percentage of dissociated acid, A^-, will be $(100 - x)\%$. So we can write that

$$\frac{100 - x}{x} = 2.04$$

This can be rearranged and solved for x:

$$\frac{100 - x}{x} = 2.04$$
$$100 - x = 2.04x$$
$$100 = 2.04x + x$$
$$100 = 3.04x$$
$$x = 32.9$$

As we set x equal to the percentage of undissociated acid, we have our answer:

32.9%