

Buffer solutions: teacher notes

Education in Chemistry

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Teacher notes

The problems that follow are based on the information in the fact sheet and allow learners to apply their mathematical skills in these different contexts.

Part 1 involves the use of logarithms to convert between $[H^+]$ and pH and between K_a and pK_a . If learners are not taking higher level courses in mathematics, they may not have met this type of calculation elsewhere, so consolidation of this skill is a good idea before the more complicated buffer solution calculations are tackled.

Part 2 gives an opportunity to practise using the Henderson–Hasselbalch equation to process buffer solution data with the aim of helping learners become more familiar with the equation and its inverse.

Part 3 has two more extended questions combining buffer solution calculations with stoichiometry in the context of making buffers by partial neutralisation of the weak acid.

Parts 1 and 2 should be accessible to learners who have studied pH, weak acids and buffer solutions, typically in the second year of a post-16 chemistry course. Part 3 is aimed at more able students at the same stage of the course who need practice at more unstructured synoptic questions.

Answers

These questions are based on the information about the lateral flow test, pH balanced shampoo and ocean acidification given on the fact sheet and will allow you to practise the mathematical skills required for calculations involving pH and buffer solutions.

Part 1: pH and pK_a calculations

1. $[H^+] = 10^{-7.40} = 3.98 \times 10^{-8} \text{ mol dm}^{-3}$

2. $\text{pH} = -\log(1.60 \times 10^{-7}) = 6.80$

3. $K_a = 10^{-10.32} = 4.79 \times 10^{-11} \text{ mol dm}^{-3}$

4. $pK_a = -\log(1.48 \times 10^{-4}) = 3.83$

3.13, the K_a for citric acid, is less than 3.83, therefore **citric acid is the stronger acid.**

5. At pH 8.2, $[H^+] = 10^{-8.2} = 6.3 \times 10^{-9} \text{ mol dm}^{-3}$
At pH 8.1, $[H^+] = 10^{-8.1} = 7.9 \times 10^{-9} \text{ mol dm}^{-3}$

Therefore the % increase = $(100 \times 7.9/6.3) - 100\% = 25\%$

Part 2: using the Henderson–Hasselbalch equation

6. Substituting the pK_a into the Henderson–Hasselbalch equation for the target pH of 7.40 gives:

$$7.40 = 7.21 + \log_{10} \frac{[HPO_4^{2-}]}{[H_2PO_4^-]}, \text{ which can be rearranged to give: } \frac{[HPO_4^{2-}]}{[H_2PO_4^-]} = 10^{7.40 - 7.21}, \text{ resulting in a ratio of } 10^{0.19} = \mathbf{1.54}.$$

Note that you can use any actual concentrations, as long as the acid/salt concentration ratio is correct. However, if the concentrations of acid and salt are too low, then the buffer will be easily overwhelmed by addition of significant amounts of acid or alkali.

7. Substituting into the Henderson–Hasselbalch equation using the pK_a for the first dissociation gives:

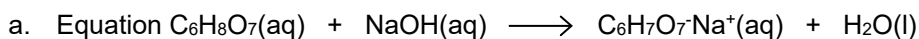
$$pH = 3.13 + \log_{10} \frac{0.0090}{0.0010}, \text{ which gives: } pH = 3.13 + 0.95 = \mathbf{4.08}$$

8. Substituting into the Henderson–Hasselbalch equation gives: $8.10 = 10.32 + \log_{10} \frac{[CO_3^{2-}]}{0.0018}$

$$\text{Rearranging this gives: } = 10^{8.10 - 10.32}, \text{ so } [CO_3^{2-}] = 0.0018 \times 10^{-2.22} = \mathbf{1.08 \times 10^{-5} \text{ mol dm}^{-3}}$$

Part 3: making buffer solutions by partial neutralisation

9.



$$\text{Amount of citric acid} = 0.100 \text{ dm}^3 \times 0.0100 \text{ mol dm}^{-3} = 0.00100 \text{ mol}$$

$$\text{Amount of sodium hydroxide added} = 0.020 \text{ dm}^3 \times 0.040 \text{ mol dm}^{-3} = 0.00080 \text{ mol}$$

Equation ratio citric acid:sodium hydroxide is 1:1, so the amount of citric acid that reacts with the sodium hydroxide = 0.00080 mol, which would give 0.00080 mol of sodium monocation salt

$$\text{Amount of citric acid left over} = 0.00100 - 0.00080 = 0.00020 \text{ mol}$$

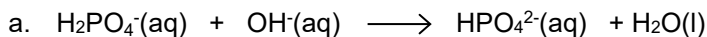
$$\text{Salt/acid concentration ratio} = 0.00080/0.00020 = \mathbf{4.0}$$

Note that although it's the salt/acid concentration ratio that is needed, it's OK to calculate the salt/acid mole ratio, as this will be the same value because both salt and acid are dissolved in the same volume of water.

- b. Substituting into the Henderson–Hasselbalch equation gives:

$$pH = 3.13 + \log_{10} 4.0, \text{ therefore: } pH = 3.13 + 0.60 = \mathbf{3.73}$$

10.



b. Amount of sodium dihydrogen phosphate = $0.500 \text{ dm}^3 \times 0.100 \text{ mol dm}^{-3} = 0.0500 \text{ mol}$

If x mol of sodium hydroxide is added and the equation ratio in part a. is 1:1

then the amount of hydrogen phosphate salt formed = x mol

and the amount of dihydrogen phosphate remaining = $0.0500 - x$ mol

The required salt/acid ratio of 1.5 = $x / (0.0500 - x)$, which rearranged gives: $0.075 = 2.5x$, so $x = 0.030$ mol sodium hydroxide

Therefore the volume of $0.200 \text{ mol dm}^{-3}$ aqueous sodium hydroxide required = $0.030 / 0.200 = \mathbf{0.15 \text{ dm}^3}$