

AS and A LEVEL CHEMISTRY B (SALTERS)

pH calculations worksheet

Instructions and answers for teachers

These instructions should accompany the OCR resource 'pH calculations worksheet' which supports OCR A Level Chemistry B.

The thumbnail shows the worksheet titled 'AS and A LEVEL CHEMISTRY B (SALTERS) pH calculations worksheet'. It contains 10 questions of increasing difficulty, covering topics like calculating pH from concentration, concentration from pH, buffer solutions, and the effect of dilution on pH. The OCR logo is visible in the bottom right corner of the document.

The Activity:

This is a series of 10 questions on pH covering all the areas in the specification in increasing depth.



This activity offers an opportunity for maths skills development.

Learning outcomes:

This lesson element relates to the specification learning outcomes O(l), O(m).

Associated materials:

'pH calculations worksheet' Lesson Element learner activity sheet.



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Introduction

This series of questions covers a range of pH calculations. This worksheet could be offered to learners as a follow-up exercise during the teaching of *Oceans* to check their understanding of the topic. Alternatively, this worksheet can be used in exam preparation.

Instructions

Encourage learners to make a sensible estimate of the answer before attempting the calculation.

Worked answers

1. HCl is a strong acid, so $[\text{H}^+] = 0.005 \text{ mol dm}^{-3}$
 $\text{pH} = -\log 0.005 = 2.3$
2. H_2SO_4 is diprotic, so each molecule gives 2 H^+ .
So $[\text{H}^+] = 4 \times 10^{-4} \text{ mol dm}^{-3}$
 $\text{pH} = -\log (4 \times 10^{-4}) = 3.4$
3. $[\text{H}^+] = 10^{-1.3} = 0.05 \text{ mol dm}^{-3}$
4. $[\text{H}^+] = \frac{1 \times 10^{-14}}{[\text{OH}^-]} = \frac{1 \times 10^{-14}}{0.0002} = 5 \times 10^{-11} \text{ mol dm}^{-3}$
 $\text{pH} = -\log (5 \times 10^{-11}) = 10.3$
5. $n(\text{H}^+) = 30 \times 1/1000 = 0.03 \text{ mol}$ $n(\text{OH}^-) = 20 \times 1/1000 = 0.02 \text{ mol}$
 $0.03 - 0.02 = 0.01 \text{ mol H}^+$ remains in the final solution
 $[\text{H}^+] = 0.01/50 \times 1000 = 0.2 \text{ mol dm}^{-3}$
 $\text{pH} = -\log 0.2 = 0.7$
6. For a weak acid, $[\text{H}^+] = \sqrt{K_a \times [\text{acid}]}$
 $\sqrt{2 \times 10^{-5} \times 0.03} = 7.7 \times 10^{-4} \text{ mol dm}^{-3}$
 $\text{pH} = 3.1$

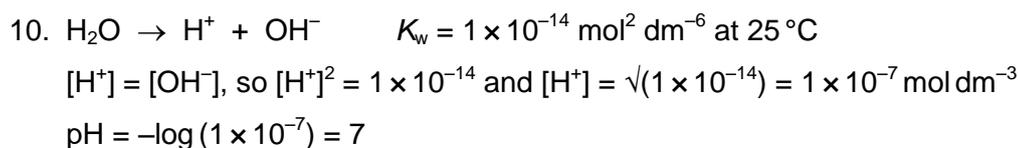


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7. $1 \text{ mol dm}^{-3} \text{ HCl}$ has a pH of zero. A more concentrated solution of HCl will be more acidic.
e.g. $2 \text{ mol dm}^{-3} \text{ HCl}$ has $[\text{H}^+] = 2 \text{ mol dm}^{-3}$; $\text{pH} = -\log 2 = -0.3$
So, yes, $\text{pH} < 0$ is possible.
Pure sulfuric acid (18 mol dm^{-3}) does not ionise, but you can add it very carefully to a little water to make a 15 mol dm^{-3} solution
 $[\text{H}^+] = 30 \text{ mol dm}^{-3}$; $\text{pH} = -\log 30 = -1.5$
8. For a buffer solution, $[\text{H}^+] = K_a[\text{acid}]/[\text{salt}]$
 $2 \times 10^{-5} \times 0.1/0.2 = 1 \times 10^{-5} \text{ mol dm}^{-3}$
 $\text{pH} = -\log (1 \times 10^{-5}) = 5.0$
- 9a. $n(\text{acid}) = 20 \times 2/1000 = 0.04 \text{ mol}$ $n(\text{NaOH}) = 20 \times 1/1000 = 0.02 \text{ mol}$
So, 0.02 mol sodium ethanoate is produced, and 0.02 mol ethanoic acid remains in 40 cm^3 .
 $[\text{acid}] = [\text{salt}] = 0.02/40 \times 1000 = 0.5 \text{ mol dm}^{-3}$
 $[\text{H}^+] = K_a[\text{acid}]/[\text{salt}] = 2 \times 10^{-5} \times 0.5/0.5 = 2 \times 10^{-5}$
 $\text{pH} = 4.7$
- 9b. (i) $n(\text{H}^+)$ added = $1 \times 1/1000 = 0.001 \text{ mol}$
original $[\text{H}^+]$ was negligible, so
 $[\text{H}^+] = 0.001 \text{ mol dm}^{-3}$ $\text{pH} = -\log 0.001 = 3$
The pH has dropped from 7 to 3.
- (ii) $n(\text{H}^+)$ added = $1 \times 1/1000 = 0.001 \text{ mol}$
from part (a), $n(\text{acid}) = n(\text{salt}) = 0.5 \text{ mol}$ in 1 dm^3 of the original solution
Assume that all the added H^+ combines with ethanoate to form ethanoic acid. Now, in the new solution
 $n(\text{acid}) = 0.501 \text{ mol}$ in 1 dm^3
 $n(\text{salt}) = 0.499 \text{ mol}$ in 1 dm^3
 $[\text{H}^+] = 2 \times 10^{-5} \times 0.501/0.499 = 2.008 \times 10^{-5} \text{ mol dm}^{-3}$
 $\text{pH} = -\log 2.008 = 4.697 \approx 4.7$
pH has not changed significantly.



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The ionisation is endothermic, so if you heat water, the equilibrium shifts to the right. The concentrations of $[\text{H}^+]$ and $[\text{OH}^-]$ increase. If $[\text{H}^+]$ increases, then the pH is lower.

So the pH decreases as temperature increases.



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