

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, and buffer solutions. Question 10 includes a note about the autoionization of water and its enthalpy change.

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, $[\text{H}^+] = \sqrt{K_a \times [\text{acid}]}$
 $\sqrt{2 \times 10^{-5} \times 0.03} = 7.7 \times 10^{-4} \text{ mol dm}^{-3}$
 $\text{pH} = 3.1$

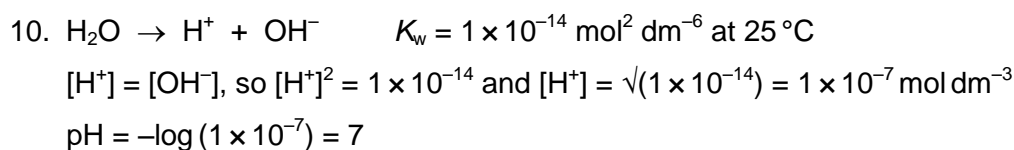


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7. $1 \text{ mol dm}^{-3} \text{ HCl}$ has a pH of zero. A more concentrated solution of HCl will be more acidic.
e.g. $2 \text{ mol dm}^{-3} \text{ HCl}$ has $[\text{H}^+] = 2 \text{ mol dm}^{-3}$; $\text{pH} = -\log 2 = -0.3$
So, yes, $\text{pH} < 0$ 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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