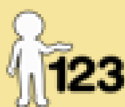
pH calculations worksheet

Instructions and answers for teachers

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

The 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.

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 $[H^+] = 0.005\ mol\ dm^{-3}$

   $pH = -\log 0.005 = 2.3$

2. H_2SO_4 is diprotic, so each molecule gives 2 H^+.

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

   $pH = -\log(4 \times 10^{-4}) = 3.4$

3. $[H^+] = 10^{-1.3} = 0.05\ mol\ dm^{-3}$

4. $[H^+] = \dfrac{1\times10^{-14}}{[OH^-]} = \dfrac{1\times10^{-14}}{0.0002} = 5 \times 10^{-11}\ mol\ dm^{-3}$

   $pH = -\log(5 \times 10^{-11}) = 10.3$

5. $n(H^+) = 30 \times 1/1000 = 0.03\ mol$ $\quad n(OH^-) = 20 \times 1/1000 = 0.02\ mol$

   $0.03 - 0.02 = 0.01\ mol\ H^+$ remains in the final solution

   $[H^+] = 0.01/50 \times 1000 = 0.2\ mol\ dm^{-3}$

   $pH = -\log 0.2 = 0.7$

6. For a weak acid, $[H^+] = \sqrt{K_a \times [acid]}$

   $\sqrt{2\times10^{-5} \times 0.03} = 7.7 \times 10^{-4}\ mol\ dm^{-3}$

   $pH = 3.1$

7. 1 mol dm^{-3} HCl has a pH of zero. A more concentrated solution of HCl will be more acidic.
e.g. 2 mol dm^{-3} HCl has [H$^+$] = 2 mol dm^{-3}; pH = –log 2 = –0.3
So, yes, 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
[H$^+$] = 30 mol dm^{-3}; pH = –log 30 = –1.5

8. For a buffer solution, [H$^+$] = K_a[acid]/[salt]
$2 \times 10^{-5} \times 0.1/0.2 = 1 \times 10^{-5}$ mol dm^{-3}
pH = –log (1×10^{-5}) = 5.0

9a. n(acid) = 20 × 2/1000 = 0.04 mol n(NaOH) = 20 × 1/1000 = 0.02 mol
So, 0.02 mol sodium ethanoate is produced, and 0.02 mol ethanoic acid remains in 40 cm^3.
[acid] = [salt] = 0.02/40 × 1000 = 0.5 mol dm^{-3}
[H$^+$] = K_a[acid]/[salt] = $2 \times 10^{-5} \times 0.5/0.5 = 2 \times 10^{-5}$
pH = 4.7

9b. (i) n(H$^+$) added = 1 × 1/1000 = 0.001 mol
original [H$^+$] was negligible, so
[H$^+$] = 0.001 mol dm^3 pH = –log 0.001 = 3
The pH has dropped from 7 to 3.

(ii) n(H$^+$) added = 1 × 1/1000 = 0.001 mol
from part (a), n(acid) = n(salt) = 0.5 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(acid) = 0.501 mol in 1 dm^3
n(salt) = 0.499 mol in 1 dm^3
[H$^+$] = $2 \times 10^{-5} \times 0.501/0.499 = 2.008 \times 10^{-5}$ mol dm^{-3}
pH = –log 2.008 = 4.697 ≈ 4.7
pH has not changed significantly.

10. $H_2O \rightarrow H^+ + OH^-$ $K_w = 1 \times 10^{-14}$ mol^2 dm^{-6} at 25 °C

$[H^+] = [OH^-]$, so $[H^+]^2 = 1 \times 10^{-14}$ and $[H^+] = \sqrt{(1 \times 10^{-14})} = 1 \times 10^{-7}$ mol dm^{-3}

pH = $-\log(1 \times 10^{-7}) = 7$

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

So the pH decreases as temperature increases.