pH, water and weak acids

OCR A H432 · Full A-level

Show your working where needed. Check the worked answers when you finish.

Name _____ Date _____

1. Nitric acid at 298 K

Calculate the pH of 0.0145 mol dm^{-3} HNO_3 at 298 K. Give your answer to two decimal places.

2. Find the conjugate pairs

$HCOOH + H_2O \rightleftharpoons HCOO^- + H_3O^+$. Identify the proton donor and proton acceptor in the forward direction. Write both conjugate acid-base pairs.

3. Barium hydroxide solution

A barium hydroxide solution contains 0.00500 mol dm^{-3} $Ba(OH)_2$. Assume complete dissociation. At this temperature Kw = 1.00×10^{-14} mol^2 dm^{-6}.

Find $[OH^-]$ and the pH.

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4. Pure water at a higher temperature

At a temperature above 298 K, Kw = 5.80×10^{-14} mol^2 dm^{-6}. Calculate the pH of pure water at this temperature. Use your result and the ion concentrations to decide whether the water is acidic, alkaline or neutral.

5. Ethanoic acid in water

Ethanoic acid has Ka = 1.80×10^{-5} mol dm^{-3}. Write the Ka expression and calculate the pH of a 0.120 mol dm^{-3} solution. State any assumptions.

6. Determine Ka from a pH reading

A 0.0800 mol dm^{-3} solution of a weak monobasic acid has pH 2.92. Calculate Ka using the weak-acid approximation, then calculate pKa.

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7. Evaluate a pH estimate

A student uses [HA] ≈ 0.00500 mol dm^{-3} for a monobasic acid with Ka = 1.00 × 10^{-3} mol dm^{-3} and starting concentration 0.00500 mol dm^{-3}. Their estimate is $[H^+]$ = 0.00224 mol dm^{-3}.

Assess the approximation using the data. Neglect water's ionisation and write the equilibrium concentration of HA in terms of x = $[H^+]$.

8. Hydrochloric acid and sodium hydroxide

Mix 25.0 cm^3 of 0.144 mol dm^{-3} HCl with 20.0 cm^3 of 0.150 mol dm^{-3} NaOH. Assume the volumes are additive.

Calculate the pH after mixing and identify the reagent in excess.

9. Extra challenge · What changes when a weak acid is diluted?

A 0.0400 mol dm^{-3} solution of a weak monobasic acid has pH 2.90. A second sample of the same solution is diluted to five times its original volume at the same temperature.

Use the weak-acid approximation to calculate Ka and the pH after dilution. Use the percentage ionisation of each sample to compare how well the approximation applies.

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Worked answers

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1. Nitric acid at 298 K

Nitric acid dissociates essentially completely, releasing one H^+ per acid molecule. Thus $[H^+] = 0.0145$ mol dm^{-3}.

$pH = -\log_{10}(0.0145) = 1.84$.

2. Find the conjugate pairs

Methanoic acid donates H^+ and water accepts H^+.

The acid-base pairs are $HCOOH/HCOO^-$ and H_3O^+/H_2O. Each pair differs by a proton.

3. Barium hydroxide solution

Each $Ba(OH)_2$ gives two OH^-, so $[OH^-] = 0.0100$ mol dm^{-3}. $[H^+] = Kw/[OH^-] = 1.00 \times 10^{-12}$ mol dm^{-3}.

pH = 12.00.

4. Pure water at a higher temperature

$[H^+] = [OH^-]$ in pure water, so $[H^+] = \sqrt{Kw} = 2.41 \times 10^{-7}$ mol dm^{-3} and pH = 6.62.

The concentrations are equal, so the water is neutral. Its neutral pH at this temperature is 6.62.

5. Ethanoic acid in water

$Ka = [H^+][CH_3COO^-]/[CH_3COOH]$. Assume $[CH_3COO^-] \approx [H^+]$ and $[CH_3COOH] \approx 0.120$ mol dm^{-3}.

$[H^+] \approx \sqrt{(1.80 \times 10^{-5} \times 0.120)} = 1.47 \times 10^{-3}$ mol dm^{-3}. pH ≈ 2.83.

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Worked answers

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6. Determine Ka from a pH reading

$[H^+]$ is 10 to the power −2.92, giving 1.202… × 10^{-3} mol dm^{-3}.

Ka ≈ $[H^+]^2/0.0800$ = 1.81 × 10^{-5} mol dm^{-3}, and pKa ≈ 4.74.

7. Evaluate a pH estimate

The estimate would require about 0.00224 ÷ 0.00500 × 100 = 45% dissociation. At that level the remaining acid concentration differs substantially from the starting value.

If $[H^+]$ is x mol dm^{-3}, x mol dm^{-3} of HA has dissociated, so [HA] = (0.00500 − x) mol dm^{-3}.

8. Hydrochloric acid and sodium hydroxide

$n(H^+)$ = 0.0250 × 0.144 = 0.00360 mol. $n(OH^-)$ = 0.0200 × 0.150 = 0.00300 mol. The acid is in excess because the ions react in a 1:1 ratio.

Excess H^+ = 0.000600 mol in 0.0450 dm^3. $[H^+]$ = 0.0133… mol dm^{-3}, so pH = 1.88.

9. Extra challenge · What changes when a weak acid is diluted?

Initially $[H^+]$ is 10 to the power −2.90, or 1.26 × 10^{-3} mol dm^{-3}. Ka ≈ $[H^+]^2/0.0400$ = 3.96 × 10^{-5} mol dm^{-3}.

After dilution [HA] starts at 0.00800 mol dm^{-3}. Approximate $[H^+]$ as $\sqrt{(3.96 \times 10^{-5} \times 0.00800)}$ = 5.63 × 10^{-4} mol dm^{-3}. This gives pH approximately 3.25.

Percentage ionisation rises from approximately 3.15% to 7.04%. The approximation [HA] at equilibrium ≈ initial [HA] is less reliable in the diluted sample.

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