

Strong and weak acids, pH and buffer solutions

Chemistry grades 11-12 (A Level · IB · Class 11-12) · Worksheet answers

Each answer shows the equation, the substitution and the result with its unit. Check your logarithms with a calculator and compare the method as well as the final value, because small rounding differences in the last digit are acceptable.

Level: Chemistry grades 11-12 (A Level · IB · Class 11-12) · **Curriculum:** IB R3.1

Problem 1 Conjugate pairs in ammonia

1. A Brønsted–Lowry base is a **proton (H^+) acceptor**.
2. NH_3 accepts H^+ from water, so **NH_3 is the base and H_2O is the acid**.
3. Pair 1: **NH_3 / NH_4^+** (base / conjugate acid). Pair 2: **H_2O / OH^-** (acid / conjugate base).

Problem 2 pH of dilute hydrochloric acid

1. $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$ (complete, strong acid).
2. $[\text{H}^+] = \mathbf{0.0250 \text{ mol/dm}^3}$.
3. $\text{pH} = -\log(0.0250) = \mathbf{1.60}$

Problem 3 Hydrogen ion concentration from pH

1. $[\text{H}^+] = 10^{-4.60} = \mathbf{2.5 \times 10^{-5} \text{ mol/dm}^3}$ (2.51×10^{-5}).
2. Neutral water: $[\text{H}^+] = 1.0 \times 10^{-7}$. Ratio = $2.51 \times 10^{-5} \div 1.0 \times 10^{-7} = \mathbf{\text{about 250 times larger}}$ ($10^{2.4} = 251$).

Problem 4 pOH and pH of sodium hydroxide

1. NaOH is a strong base, so $[\text{OH}^-] = \mathbf{0.0040 \text{ mol/dm}^3}$.
2. $\text{pOH} = -\log(0.0040) = \mathbf{2.40}$.
3. $\text{pH} = 14 - 2.40 = \mathbf{11.60}$

Problem 5 Comparing acid strength by pKa

1. $\text{pKa} = -\log(6.3 \times 10^{-5}) = \mathbf{4.20}$.
2. **Chloroethanoic acid** is stronger: its pKa is lower, so its K_a is larger and it ionises more.
3. $K_a = 10^{-2.85} = \mathbf{1.4 \times 10^{-3} \text{ mol/dm}^3}$.

Problem 6 pH of methanoic acid

1. $K_a = [\text{H}^+][\text{HCOO}^-] / [\text{HCOOH}]$.
2. $[\text{H}^+] = \sqrt{(1.6 \times 10^{-4} \times 0.150)} = \sqrt{(2.4 \times 10^{-5})} = \mathbf{4.9 \times 10^{-3} \text{ mol/dm}^3}$.
3. $\text{pH} = -\log(4.9 \times 10^{-3}) = \mathbf{2.31}$.
4. Fraction ionised $= 4.9 \times 10^{-3} \div 0.150 = 3.3 \%$, below 5 %, so $[\text{HA}] \approx c$ is acceptable.

Problem 7 Percentage ionisation of ethanoic acid

1. $[\text{H}^+] = \sqrt{(1.8 \times 10^{-5} \times 0.0500)} = \sqrt{(9.0 \times 10^{-7})} = \mathbf{9.5 \times 10^{-4} \text{ mol/dm}^3}$.
2. % ionised $= 9.49 \times 10^{-4} \div 0.0500 \times 100 = \mathbf{1.9 \%}$.
3. It **rises** (to about 6 %). Dilution shifts $\text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+$ towards the side with more particles, which is the ions.

Problem 8 Diluting nitric acid

1. $n(\text{H}^+) = 0.0100 \text{ dm}^3 \times 1.00 \text{ mol/dm}^3 = 0.0100 \text{ mol}$; $c = 0.0100 \div 0.500 = \mathbf{0.0200 \text{ mol/dm}^3}$.
2. $\text{pH} = -\log(0.0200) = \mathbf{1.70}$.
3. Water added $= 500 - 10.0 = \mathbf{490 \text{ cm}^3}$.

Problem 9 Excess acid after partial neutralisation

1. $n(\text{H}^+) = 0.0300 \times 0.120 = 3.60 \times 10^{-3} \text{ mol}$; $n(\text{OH}^-) = 0.0200 \times 0.100 = 2.00 \times 10^{-3} \text{ mol}$.
2. Excess $\text{H}^+ = 3.60 \times 10^{-3} - 2.00 \times 10^{-3} = \mathbf{1.60 \times 10^{-3} \text{ mol}}$.
3. Total volume 50.0 cm^3 : $[\text{H}^+] = 1.60 \times 10^{-3} \div 0.0500 = 0.0320 \text{ mol/dm}^3$; $\text{pH} = \mathbf{1.49}$.

Problem 10 K_a from a measured pH

1. $[\text{H}^+] = 10^{-2.85} = \mathbf{1.41 \times 10^{-3} \text{ mol/dm}^3}$.
2. $[\text{A}^-] = [\text{H}^+]$; $K_a = (1.41 \times 10^{-3})^2 \div 0.200 = \mathbf{1.0 \times 10^{-5} \text{ mol/dm}^3}$.
3. $\text{p}K_a = \mathbf{5.00}$.

Problem 11 A lactic acid buffer

1. $\text{pH} = \text{p}K_a + \log([\text{A}^-]/[\text{HA}])$.
2. $\text{pH} = 3.85 + \log(0.15 \div 0.30) = 3.85 - 0.30 = \mathbf{3.55}$.

Problem 12 How a buffer neutralises added acid

1. $\text{CH}_3\text{COO}^- + \text{H}^+ \rightarrow \text{CH}_3\text{COOH}$.
2. One A^- is used up for each H^+ : HA rises from **4 to 5** and A^- falls from **3 to 2**.
3. The added ion is removed by the conjugate partner, so the ratio $[\text{A}^-]/[\text{HA}]$ changes but both species stay plentiful; $\log(\text{ratio})$ moves very little, so the pH hardly changes.

Problem 13 Ionic product of water at 50 °C

1. $[H^+] = [OH^-] = \sqrt{5.5 \times 10^{-14}} = 2.3 \times 10^{-7} \text{ mol/dm}^3$.
2. $pH = -\log(2.35 \times 10^{-7}) = 6.63$.
3. **Neutral**, because $[H^+] = [OH^-]$. Neutral pH is 7 only at 25 °C; ionisation of water is endothermic, so K_w grows with temperature.

Problem 14 pH of aqueous ammonia

1. $[OH^-] = \sqrt{(1.8 \times 10^{-5} \times 0.100)} = 1.34 \times 10^{-3} \text{ mol/dm}^3$.
2. $pOH = -\log(1.34 \times 10^{-3}) = 2.87$.
3. $pH = 14.00 - 2.87 = 11.13$

Problem 15 Reading pKa from a titration curve

1. The steep rise is centred at **24.0 cm³**.
2. At 12.0 cm³ the pH is **4.87** (half-equivalence, so $pH = pK_a$).
3. pK_a 4.87 matches **propanoic acid**; $K_a = 10^{-4.87} = 1.35 \times 10^{-5} \text{ mol/dm}^3$.

Problem 16 Equivalence point and indicator

1. Equivalence volume 25.0 cm³; $n = 2.50 \times 10^{-3} \text{ mol}$ in 50.0 cm³; $c = 0.0500 \text{ mol/dm}^3$.
2. $K_b = 1.0 \times 10^{-14} \div 1.8 \times 10^{-5} = 5.6 \times 10^{-10}$. $[OH^-] = \sqrt{(5.6 \times 10^{-10} \times 0.0500)} = 5.3 \times 10^{-6} \text{ mol/dm}^3$; $pOH = 5.28$; $pH = 8.72$.
3. **Phenolphthalein**: its range (8.2–10.0) contains the equivalence pH and lies on the steep part of the curve. Methyl orange changes in the buffer region, far too early.

Problem 17 Preparing an ammonia buffer

1. $9.40 = 9.25 + \log(\text{ratio})$, so $\log(\text{ratio}) = 0.15$ and $\text{ratio} = 10^{0.15} = 1.41$.
2. $n(NH_4^+) = 0.100 \times 0.200 = 0.0200 \text{ mol}$; $n(NH_3) = 1.41 \times 0.0200 = 0.0283 \text{ mol}$.
3. $V = 0.0283 \text{ mol} \div 1.00 \text{ mol/dm}^3 = 0.0283 \text{ dm}^3 = 28.3 \text{ cm}^3$, then make up to 250 cm³ with water.

Problem 18 Adding acid to a buffer and to water

1. Equal concentrations, so $pH = pK_a = 4.74$.
2. Start: 0.0500 mol each. HCl converts A^- to HA: HA = 0.0600 mol, $A^- = 0.0400 \text{ mol}$. $pH = 4.74 + \log(0.0400 \div 0.0600) = 4.74 - 0.18 = 4.56$ (4.564).
3. $[H^+] = 0.0100 \div 0.500 = 0.0200 \text{ mol/dm}^3$; $pH = 1.70$.
4. The buffer changes by about **0.18** units; the water changes by about **5.3** units.

Problem 19 Mixing weak acid and alkali

1. $n(HA) = 0.0400 \times 0.150 = 6.00 \times 10^{-3} \text{ mol}$; $n(OH^-) = 0.0100 \times 0.200 = 2.00 \times 10^{-3} \text{ mol}$.
2. $HA + OH^- \rightarrow A^- + H_2O$: HA = $4.00 \times 10^{-3} \text{ mol}$, $A^- = 2.00 \times 10^{-3} \text{ mol}$.

3. $\text{pH} = 4.20 + \log(2.00 \div 4.00) = 4.20 - 0.30 = \mathbf{3.90}$.

4. $V = 6.00 \times 10^{-3} \div 0.200 = 0.0300 \text{ dm}^3 = \mathbf{30.0 \text{ cm}^3}$.

Problem 20 A very dilute acid

1. pH 8.00 would be alkaline, but adding an acid to water cannot make it alkaline: **water's own ionisation (10^{-7} mol/dm^3) is far bigger** than the acid's contribution and must be included.

2. $x = 1.0 \times 10^{-8} + 1.0 \times 10^{-14}/x$, so $x^2 - 1.0 \times 10^{-8}x - 1.0 \times 10^{-14} = 0$. $x = (1.0 \times 10^{-8} + \sqrt{(4.01 \times 10^{-14})}) \div 2 = 1.05 \times 10^{-7} \text{ mol/dm}^3$; pH = **6.98**.