pH calculations for weak acids

How can we calculate the pH of a weak acid? How do we know $[H_3O^+]$ if the solution has not fully dissociated?

We use a new constant!! K_a

$HA + H_2O \rightleftharpoons H_3O^+ + A^-$

$$K_a = \frac{[H_3O^+]\,[A^-]}{[HA]}$$

The higher the K_a the stronger the acid.

$pK_a = -\log_{10}K_a$ $\qquad K_a = 10^{-pKa}$

The lower the pK_a the stronger the acid.

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2014 Exam Q1 b

(b) Hypochlorous acid has a pK_a of 7.53. Another weak acid, hydrofluoric acid, HF, has a pK_a of 3.17.

A 0.100 mol L^{-1} solution of each acid was prepared by dissolving it in water.

Compare the pHs of these two solutions.

No calculations are necessary.

(b)	Hydrofluoric acid is a stronger acid/more acidic/dissociates more because it has a smaller pK_a (larger K_a) than hypochlorous acid. So HF will therefore have a higher $[H_3O^+]$. As $[H_3O^+]$ increases, the pH decreases, so HF will have a lower pH than HOCl. (pH HF = 2.09, HOCl = 4.27)

pH calculations for weak acids

We make two assumptions to calculate the pH of a weak acid.

1. $[H_3O^+] = [A^-]$ (no H_3O^+ from the water)
2. $[HA]$ = starting concentration of acid (no dissociation at all)

$HA + H_2O \rightleftharpoons H_3O^+ + A^-$ $\quad K_a(HA) = \frac{[H_3O^+][A^-]}{[HA]}$ $\quad K_a = \frac{[H_3O^+]^2}{[HA]}$

For example: Find the pH of a 1.50 $mol.L^{-1}$ solution of HOBr. $K_a(HOBr) = 2.40 \times 10^{-9.}$

$K_a = \frac{[H_3O^+]^2}{[HOBr]}$ $\quad 2.4 \times 10^{-9} = \frac{[H_3O^+]^2}{1.5}$

$HOBr + H_2O \rightleftharpoons OBr^- + H_3O^+$

$[H_3O^+] = \sqrt{(2.4 \times 10^{-9}) \times 1.5}$

$[H_3O^+] = 6 \times 10^{-5}$

$pH = -\log_{10}(6 \times 10^{-5})$

$pH = 4.22$

pH calculations for weak acids

Find the pH of 0.0500 mol.L^{-1} HF. K_a(HF) = 7.2 x 10^{-4}

Find the pH of 0.750 mol.L^{-1} CH_3COOH. pK_a(CH_3COOH) = 4.75

The pH of a solution of acetic acid is 3.42. What is the concentration of the acid solution? K_a(CH_3COOH) = 1.78 x 10^{-5}

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2013 Exam Q3

Sometimes part of a bigger question on buffers/titrations (last topics)

QUESTION THREE

20.0 mL of 0.0896 mol L^{-1} ethanoic acid is titrated with 0.100 mol L^{-1} sodium hydroxide.

$pK_a(CH_3COOH) = 4.76$

(a) Calculate the pH of the ethanoic acid before any NaOH is added.

THREE (a)

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$$

$$[H_3O^+] = \sqrt{1.74 \times 10^{-5} \times 0.0896} \text{ mol L}^{-1}$$

$$= 1.25 \times 10^{-3} \text{ mol L}^{-1}$$

$$pH = -\log[H_3O^+] = 2.90$$

Achievement	Achievement with Merit
• Correct process.	• Correct pH.

pH calculations for weak bases

Same process as weak acids – but use K_b

For $A^- + H_2O \rightleftharpoons HA + OH^-$ $\quad K_b = \dfrac{[HA]\,[OH^-]}{[A^-]}$

$K_w = K_a \times K_b$ if question gives you K_a of conjugate acid

Steps:

find K_b from K_a and K_w

find $[OH^-]$ from K_b

then calculate $[H_3O^+]$ then pH

OR

calculate pOH and convert to pH.

pH calculations for weak bases

For example: Find the pH of a 0.53 mol.L^{-1} solution of sodium ethanoate ($NaCH_3COO$).

$K_a(CH_3COOH) = 1.74 \times 10^{-5}$.

$NaCH_3COO \rightarrow Na^+ + CH_3COO^-$

$CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$

$$K_b = \frac{K_w}{K_a} = \frac{10^{-14}}{1.74 \times 10^{-5}}$$

$K_b = 5.75 \times 10^{-10}$

$$K_b = \frac{[OH^-]^2}{[CH_3COO^-]}$$

$$5.75 \times 10^{-10} = \frac{[OH^-]^2}{0.53}$$

$$[OH^-] = \sqrt{(5.75 \times 10^{-10}) \times 0.53}$$

$[OH^-] = 1.75 \times 10^{-5}$

$pOH = -\log_{10}(1.75 \times 10^{-5})$

$pOH = 4.76$

$pH = 14 - pOH$

$pH = 9.24$

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Workbook pg 234 – 235 if finished

2012 Exam Q3 a

QUESTION THREE

(a) Calculate the pH of 0.150 mol L^{-1} aqueous ammonia, NH_3.

$pK_a\ (NH_4^+) = 9.24$

Concentration calculations from pH

You can also be asked to calculate concentration if given the pH.

Again $K_a = \frac{[H_3O^+]\,[A^-]}{[HA]}$ and $K_a = \frac{[H_3O^+]^2}{[HA]}$

Rearranging gives $[HA] = \frac{[H_3O^+]^2}{[K_a]}$

Can calculate from pH

For example: A solution of glycolic acid (HG) has a pH of 2.00. Show by calculation that the concentration is 0.675 $mol.L^{-1}$

If you are asked to do this for a base you need to calculate K_b from K_a and $[OH^-]$ from pH then carry on as normal using K_b expression.

2011 Exam Q 1 c

(c) A solution prepared by dissolving hydrogen fluoride in water has a pH of 2.34.

Calculate the concentration of the hydrogen fluoride in the solution.

pK_a (HF) = 3.17

2014 Exam Q 3 b

A titration was carried out by adding hydrobromic acid, HBr, to 20.0 mL of aqueous methylamine, CH_3NH_2, solution.

The equation for the reaction is:

$$CH_3NH_2 + HBr \rightarrow CH_3NH_3^+ + Br^-$$

$$K_a(CH_3NH_3^+) = 2.29 \times 10^{-11}$$

(b) The aqueous methylamine, CH_3NH_2, solution has a pH of 11.8 before any HBr is added.

Show by calculation that the concentration of this solution is 0.0912 mol L^{-1}.

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{10^{-14}}{10^{-11.8}}$$

$$= 6.31 \times 10^{-3} \text{ mol } L^{-1}$$

$$K_b = \frac{[OH^-]^2}{[CH_3NH_2]}$$

$$4.37 \times 10^{-4} = \frac{(6.31 \times 10^{-3})^2}{[CH_3NH_2]}$$

$$[CH_3NH_2] = \frac{(6.31 \times 10^{-3})^2}{4.37 \times 10^{-4}}$$

$$[CH_3NH_2] = 0.0912 \text{ mol } L^{-1}$$

Achievement	Achievement with Merit	Achievement with Excellence
• Calculates $[OH^-]$ / $[H_3O^+]$ / K_b • Uses suitable process with more than one error. OR Rearranges K_b / K_a expression so $[CH_3NH_2]$ is the subject.	• Correct method but an error in the calculation.	• Correct answer with a clear method.