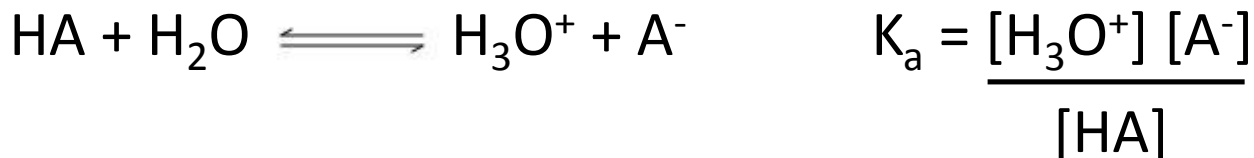


pH calculations for weak acids

How can we calculate the pH of a weak acid? How do we know $[\text{H}_3\text{O}^+]$ if the solution has not fully dissociated?

We use a new constant!! K_a



The higher the K_a the stronger the acid.

$$\text{p}K_a = -\log_{10} K_a \quad K_a = 10^{-\text{p}K_a}$$

The lower the $\text{p}K_a$ the stronger the acid.

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.

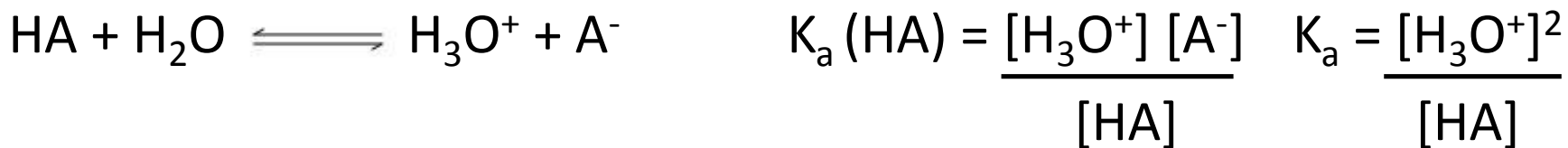
No calculations are necessary.

- (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 $[\text{H}_3\text{O}^+]$. As $[\text{H}_3\text{O}^+]$ 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. $[\text{H}_3\text{O}^+] = [\text{A}^-]$ (no H_3O^+ from the water)
2. $[\text{HA}] =$ starting concentration of acid (no dissociation at all)



For example: Find the pH of a 1.50 mol.L^{-1} solution of HOBr.

$$K_a(\text{HOBr}) = 2.40 \times 10^{-9}.$$

$$K_a = \frac{[\text{H}_3\text{O}^+]^2}{[\text{HOBr}]} \quad 2.4 \times 10^{-9} = \frac{[\text{H}_3\text{O}^+]^2}{1.5} \quad [\text{H}_3\text{O}^+] = \sqrt{(2.4 \times 10^{-9}) \times 1.5}$$
$$[\text{H}_3\text{O}^+] = 6 \times 10^{-5}$$



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

$$\text{pH} = 4.22$$

pH calculations for weak acids

Find the pH of 0.0500 mol.L⁻¹ HF. $K_a(\text{HF}) = 7.2 \times 10^{-4}$

Find the pH of 0.750 mol.L⁻¹ CH₃COOH. $pK_a(\text{CH}_3\text{COOH}) = 4.75$

The pH of a solution of acetic acid is 3.42. What is the concentration of the acid solution? $K_a(\text{CH}_3\text{COOH}) = 1.78 \times 10^{-5}$

2013 Exam Q3

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

QUESTION THREE

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

$$\text{p}K_{\text{a}} (\text{CH}_3\text{COOH}) = 4.76$$

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

THREE

(a)

$$K_{\text{a}} = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

$$[\text{H}_3\text{O}^+] = \sqrt{1.74 \times 10^{-5} \times 0.0896} \text{ mol L}^{-1}$$

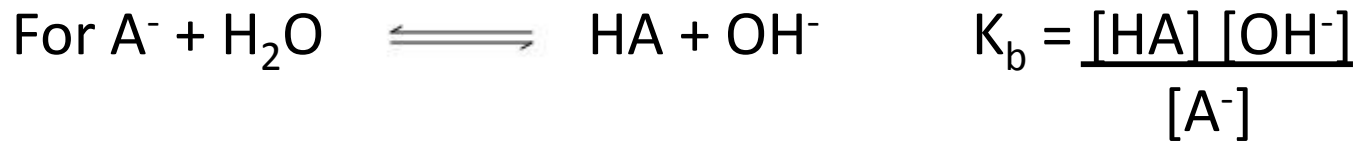
$$= 1.25 \times 10^{-3} \text{ mol L}^{-1}$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = 2.90$$

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

pH calculations for weak bases

Same process as weak acids – but use K_b



$K_w = K_a \times K_b$ 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 (NaCH₃COO).

$$K_a(\text{CH}_3\text{COOH}) = 1.74 \times 10^{-5}$$



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

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

$$K_b = \frac{[\text{OH}^-]^2}{[\text{CH}_3\text{COO}^-]}$$

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

$$\text{pH} = 14 - \text{pOH}$$

$$\text{pH} = 9.24$$

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

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

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

$$\text{pOH} = 4.76$$

Workbook pg 232

Workbook pg 234 – 235 if finished

2012 Exam Q3 a

QUESTION THREE

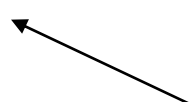
(a) Calculate the pH of 0.150 mol L^{-1} aqueous ammonia, NH_3 .

$$\text{p}K_{\text{a}}(\text{NH}_4^+) = 9.24$$

Concentration calculations from pH

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

Again
$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$
 and
$$K_a = \frac{[\text{H}_3\text{O}^+]^2}{[\text{HA}]}$$

Rearranging gives
$$[\text{HA}] = \frac{[\text{H}_3\text{O}^+]^2}{[K_a]}$$
  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 $[\text{OH}^-]$ 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.

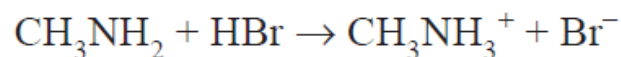
Calculate the concentration of the hydrogen fluoride in the solution.

$$pK_a(\text{HF}) = 3.17$$

2014 Exam Q 3 b

A titration was carried out by adding hydrobromic acid, HBr, to 20.0 mL of aqueous methylamine, CH₃NH₂, solution.

The equation for the reaction is:



$$K_a(\text{CH}_3\text{NH}_3^+) = 2.29 \times 10^{-11}$$

- (b) The aqueous methylamine, CH₃NH₂, 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.

$$[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{10^{-14}}{10^{-11.8}} = 6.31 \times 10^{-3} \text{ mol L}^{-1}$$

$$K_b = \frac{[\text{OH}^-]^2}{[\text{CH}_3\text{NH}_2]}$$

$$4.37 \times 10^{-4} = \frac{(6.31 \times 10^{-3})^2}{[\text{CH}_3\text{NH}_2]}$$

$$[\text{CH}_3\text{NH}_2] = \frac{(6.31 \times 10^{-3})^2}{4.37 \times 10^{-4}}$$

$$[\text{CH}_3\text{NH}_2] = 0.0912 \text{ mol L}^{-1}$$

Achievement	Achievement with Merit	Achievement with Excellence
<ul style="list-style-type: none"> Calculates $[\text{OH}^-] / [\text{H}_3\text{O}^+] / K_b$ Uses suitable process with more than one error. OR Rearranges K_b / K_a expression so $[\text{CH}_3\text{NH}_2]$ is the subject. 	<ul style="list-style-type: none"> Correct method but an error in the calculation. 	<ul style="list-style-type: none"> Correct answer with a clear method.