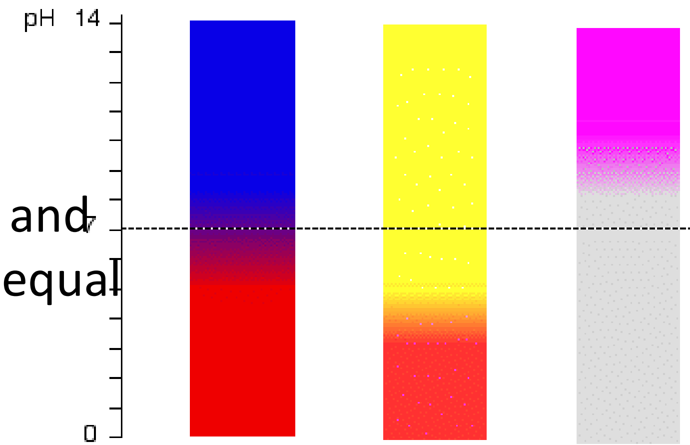
Do now:

What do you think the following words mean?

Titration curve

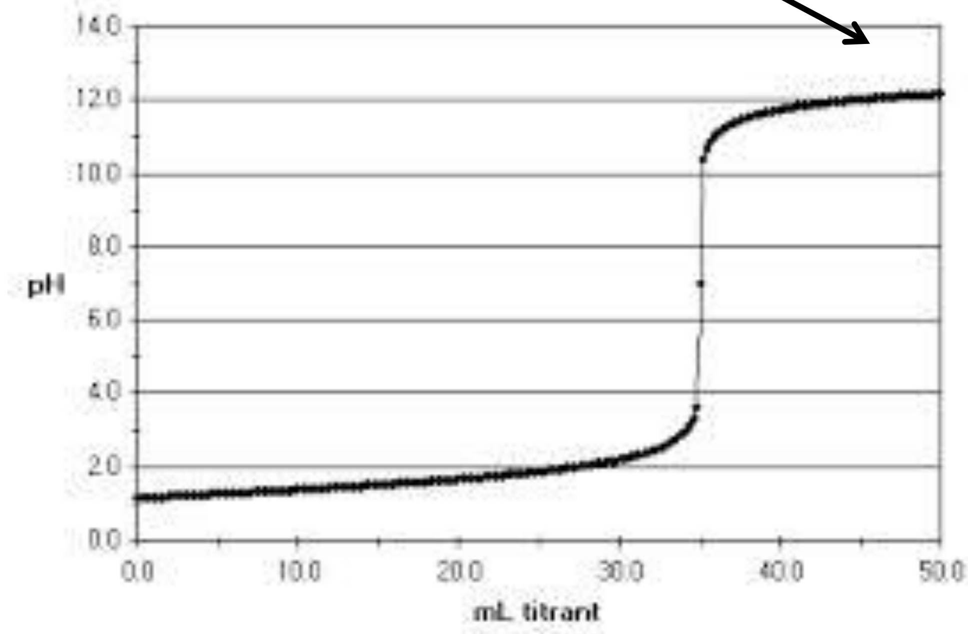
Equivalence point

Amount of acid and base present is equal



End point

Point where the indicator changes colour

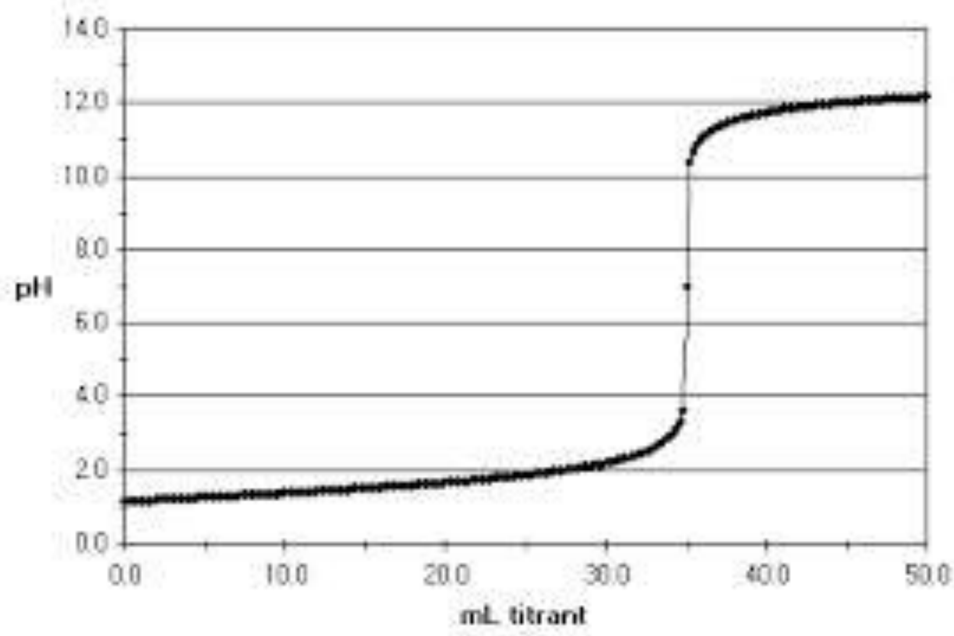


Indicator

Titrations

What do we need to be able to do?

- Discuss reasons for the pH of titrated solution at key points
- Discuss changes in concentrations of species in solution
- Calculate pH at key points on graph
- Choosing indicators

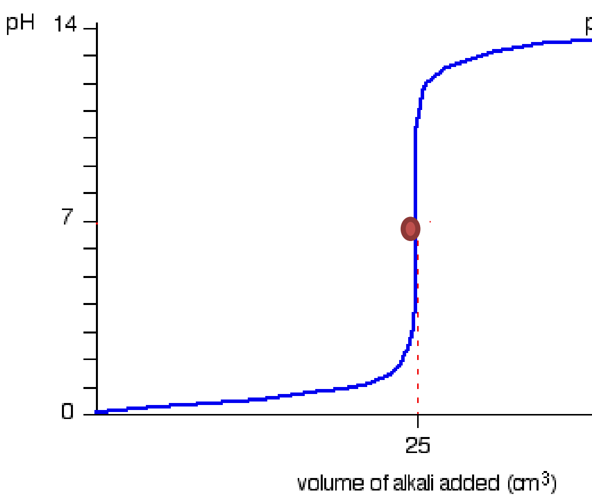


Titration curves

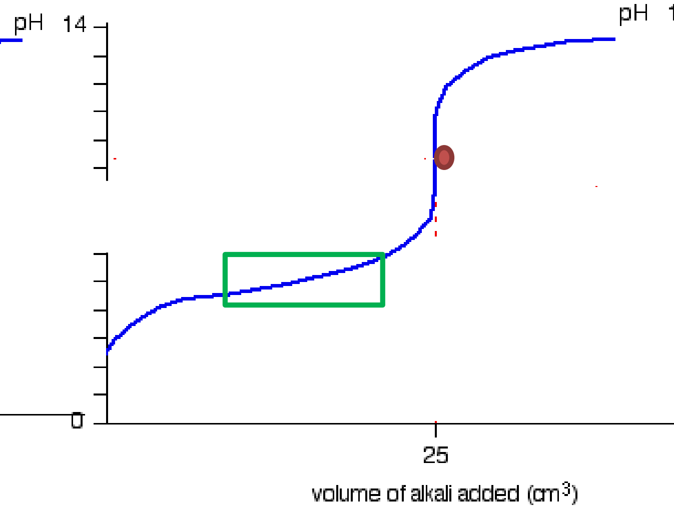
pH v volume

<u>3 types</u>

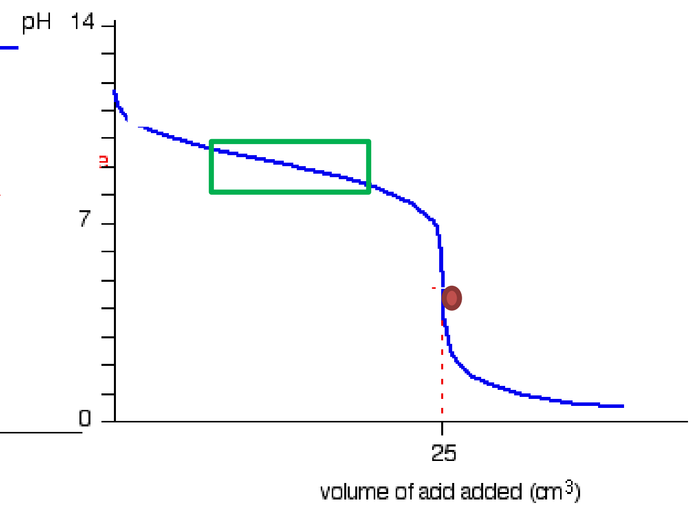
Strong acid
Strong base



Weak acid
Strong base



Strong acid
Weak base



Where is the equivalence point?

Do now:

What is the equivalence point of a titration?

Why do titrations between strong acids and weak bases OR strong bases and weak acids have buffer zones?

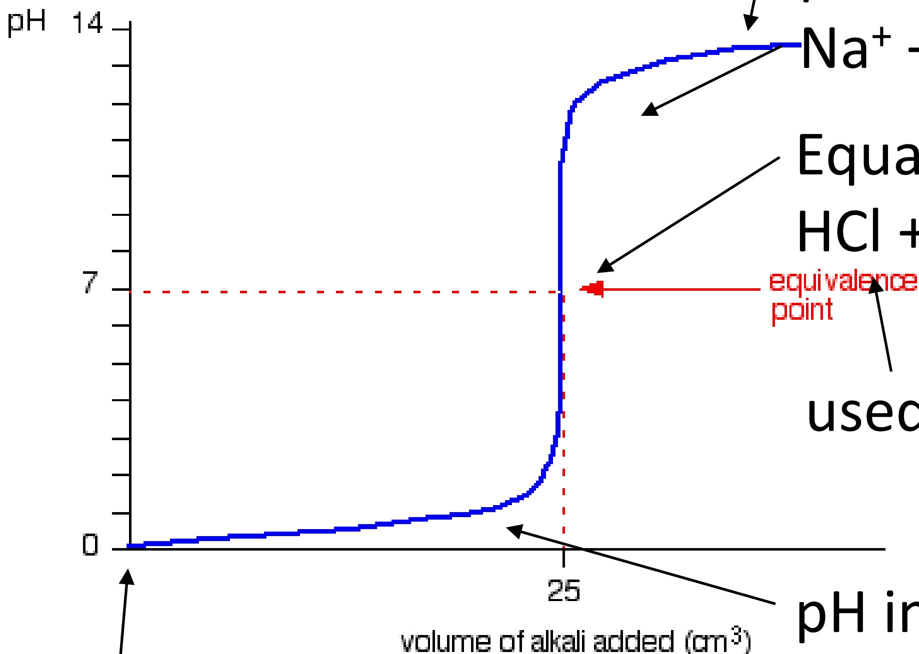
Strong acid strong base titration

eg NaOH and HCl

pH entirely dependent on [NaOH]

pH increasing

$Na^+ + Cl^- + NaOH \rightarrow 2\ Na^+ + Cl^- + OH^-$



Equal amount of HCl and NaOH

$HCl + NaOH \rightarrow Na^+ + Cl^- + H_2O$

used up

present

produced

pH increasing

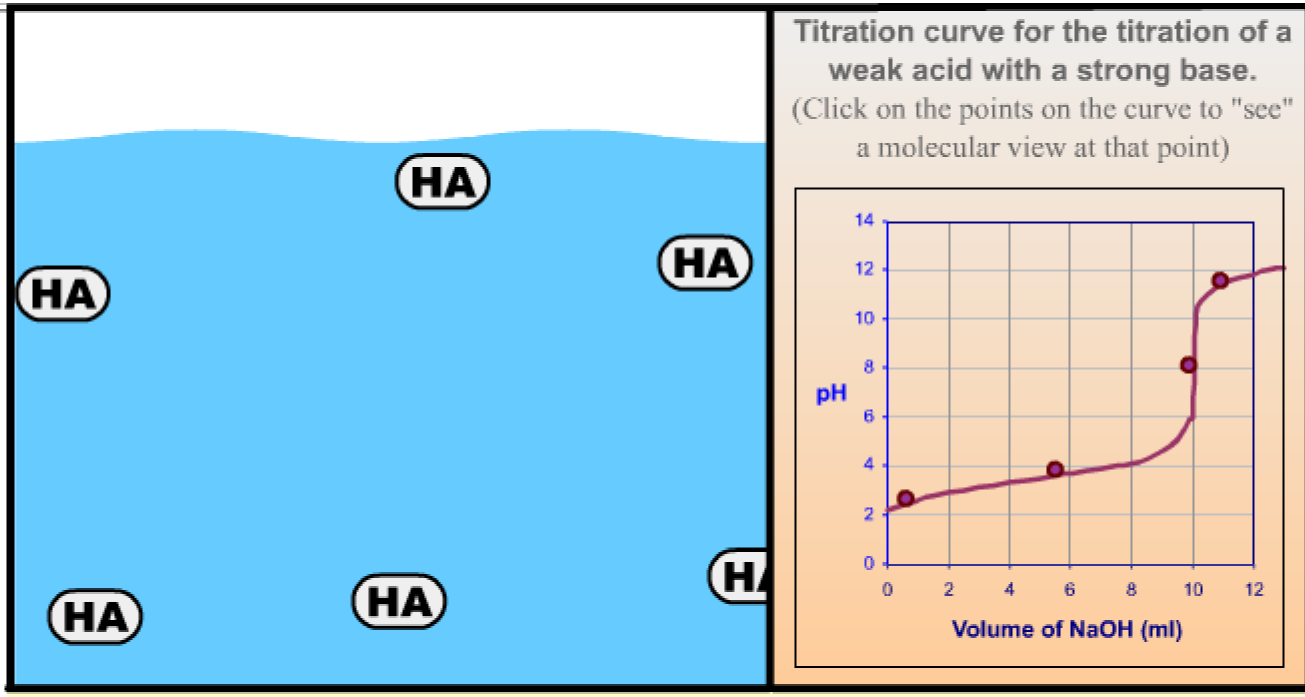
$HCl + NaOH \rightarrow Na^+ + Cl^- + H_2O$

used up by reaction with HCl

pH entirely dependent on [HCl]

[HCl] decreasing

Weak acid strong base titration



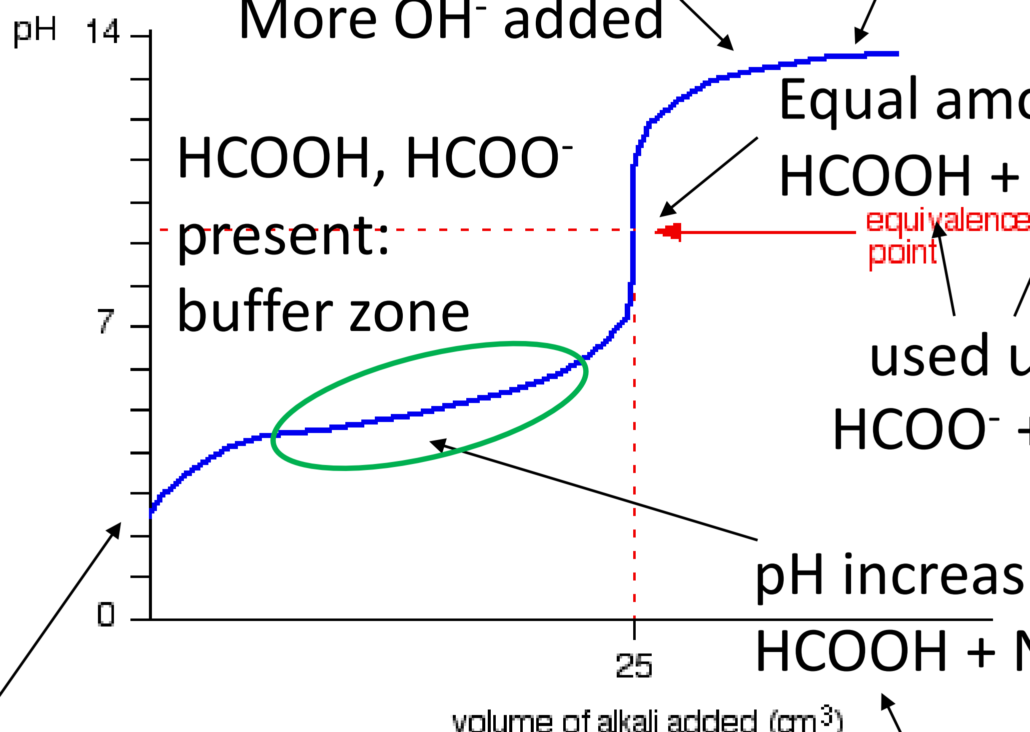
In this animation, we will show various stages of the titration of a weak acid, HA, with a strong base, NaOH.
As you may remember from the "Weak Acid" and "Buffer" animations, only a small fraction of weak acid molecules dissociates to hydronium ion, H_3O^+, and the conjugate base A^-, when dissolved in water.
In this animation we will show HA only in its non-dissociated form because it is the dominant species in solution.

Weak acid strong base titration

eg NaOH and HCOOH

pH entirely dependent on [NaOH]

pH increasing
More OH^- added



$HCOOH$, $HCOO^-$ present:
buffer zone

Equal amount of HCOOH and NaOH

$HCOOH + NaOH \rightarrow Na^+ + HCOO^- + H_2O$

used up present

$HCOO^- + H_2O \rightarrow HCOOH + OH^-$

produced

pH increasing

$HCOOH + NaOH \rightarrow Na^+ + HCOO^- + H_2O$

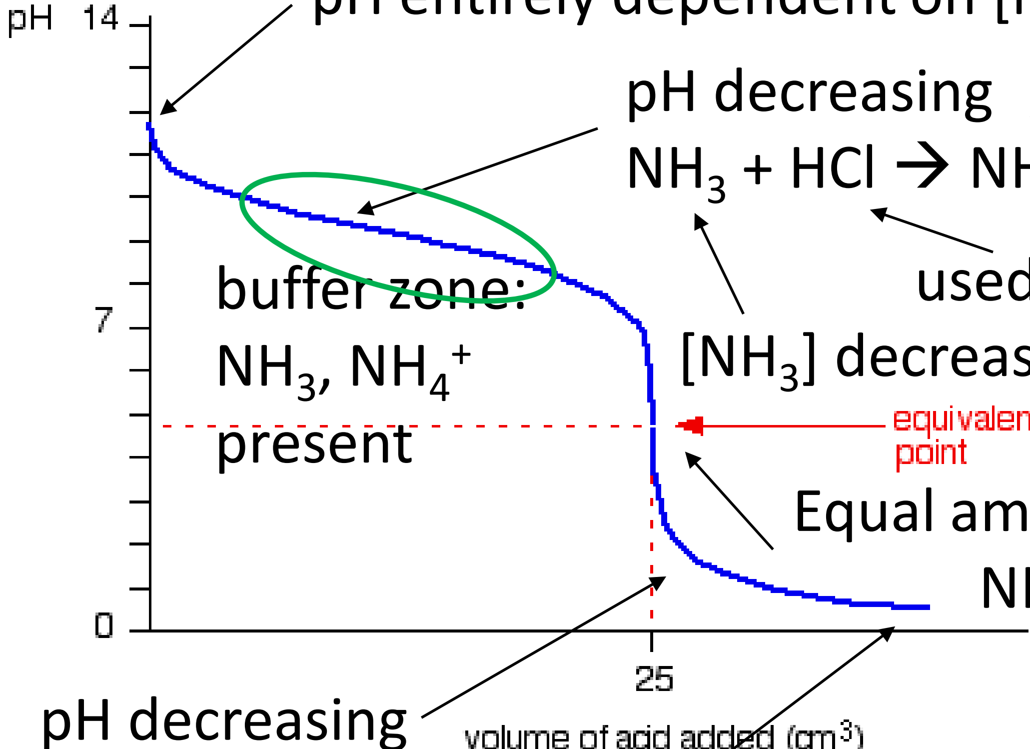
used up by reaction with HCOOH

pH entirely dependent on [HCOOH]

[HCOOH] decreasing

Strong acid weak base titration

eg NH_3 and HCl



pH entirely dependent on $[NH_3]$

pH decreasing

produced

$NH_3 + HCl \rightarrow NH_4^+ + Cl^-$

used up by reaction with NH_3

buffer zone:

NH_3, NH_4^+

present

$[NH_3]$ decreasing

Equal amount of NH_3 and HCl

$NH_3 + HCl \rightarrow NH_4^+ + Cl^-$

used up present

pH decreasing

More H_3O^+ added

$NH_4^+ + H_2O \rightarrow NH_3 + H_3O^+$

pH entirely dependent on [HCl]

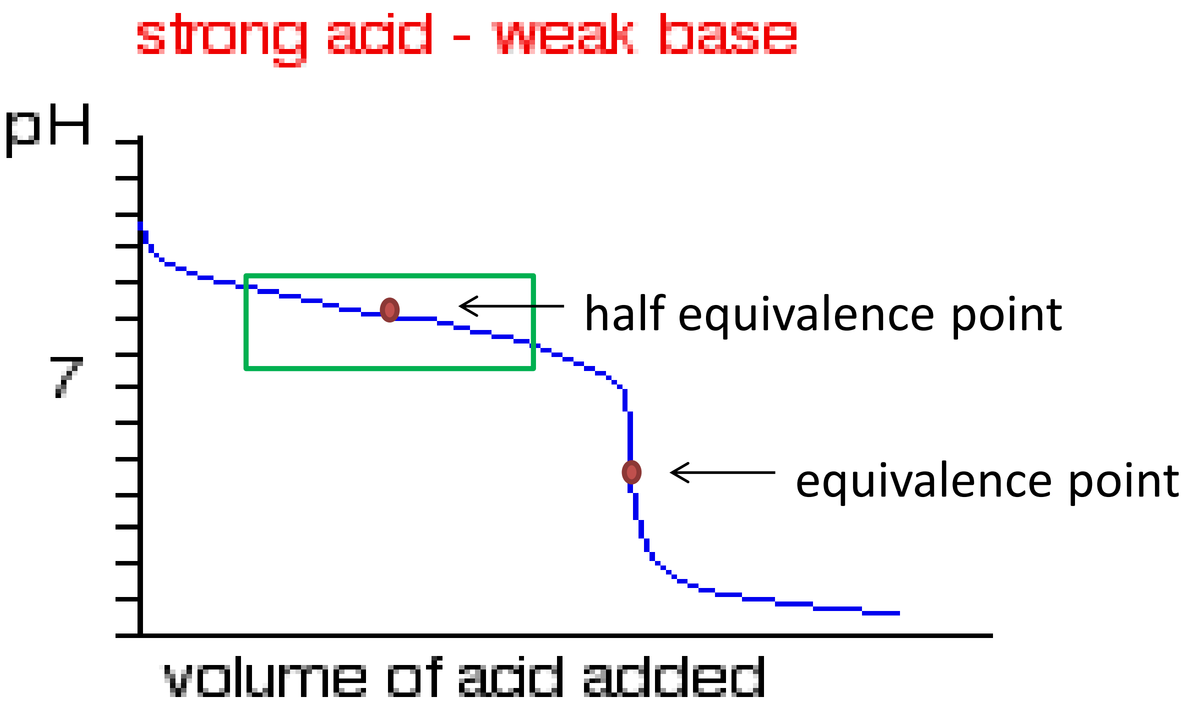
Buffer zones

Recap: What is a buffer solution?

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

$$pK_a = pH - \log_{10}\left(\frac{[A^-]}{[HA]}\right)$$

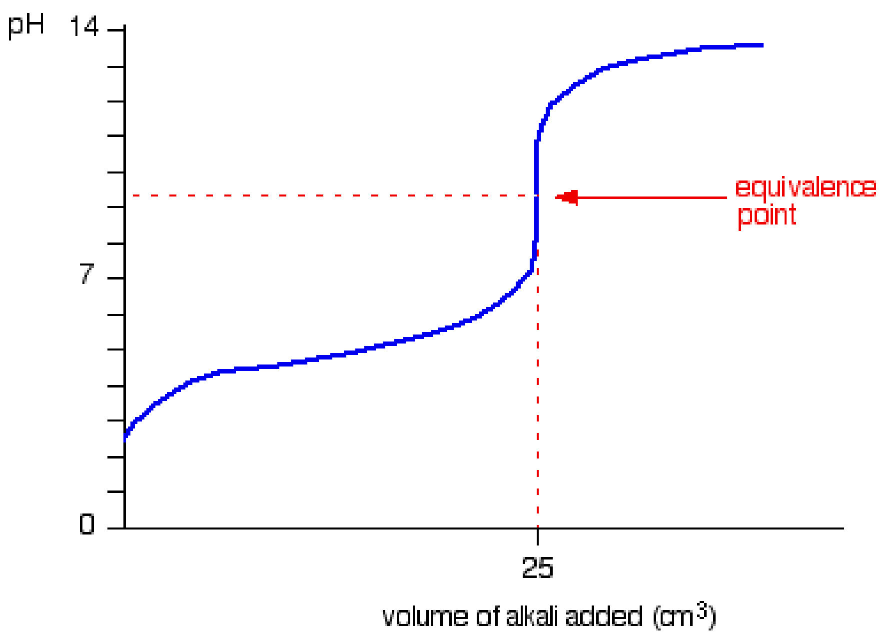
When $A^- = HA$
then $pK_a = pH$



Key points

In exams you are usually asked to:

- Discuss the buffer zone
- Discuss the equivalence point and its pH and/or calculate its pH
- Discuss the change in the concentration of species in solution as the titration progresses
- Calculate initial pH of weak acid/base solution



2013 Exam Q3 b

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

pK_a $(CH_3COOH) = 4.76$

(b) Halfway to the equivalence point of the titration, the pH = pK_a of the ethanoic acid.

Discuss the reason for this.

What would be the A answer?

- Recognises that there are equimolar quantities of ethanoic acid and sodium ethanoate.

What would be the M answer?

- Relates equation correctly to explanation.

Halfway to equivalence point, half of the ethanoic acid has been used up. There are now equimolar quantities of ethanoic acid and sodium ethanoate.

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

According to the equation when $[CH_3COOH] = [CH_3COO^-]$
then $K_a = [H_3O^+]$
So $pK_a = pH$.

(c) (i) Discuss the change in the concentration of species in solution, as the first 5.00 mL of NaOH is added to the 20.0 mL of ethanoic acid.

Your answer should include chemical equations.

No calculations are required.

What is happening when NaOH is added?
What key things is the examiner looking for?

• Correct equation *minor error.* • Correct statement relating to change in concentration of 1 species.	Correct equation and correctly describes the change in concentration of 2 species.	• Correct equation. AND Correctly describes the change in concentration of the 4 species.

$NaOH(aq) + CH_3COOH(aq) \rightarrow NaCH_3COO(aq) + H_2O(l)$ (1)

$[CH_3COO^-]$ increases as it is formed in reaction (1).

$[Na^+]$ increases as NaOH is added (1).

$[CH_3COOH]$ decreases as it reacts with NaOH (1).

$[H_3O^+]$ decreases because $[CH_3COO^-]$ / $[CH_3COOH]$ increases and K_a is a constant.

$[OH^-]$ increases because $[H_3O^+]$ decreases and $[H_3O^+]$ $[OH^-]$ is constant.

2014 Exam Q3 a

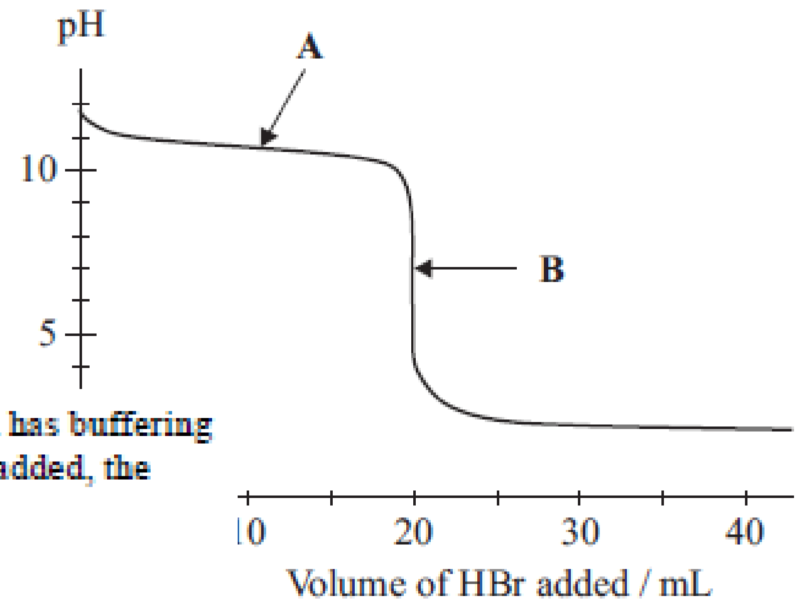
A titration was carried out by adding hydrobromic acid, HBr, to 20.0 mL of aqueous methylamine, CH_3NH_2, solution.

The equation for the reaction is:

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

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

The curve for this titration is given below:



At point A, $[CH_3NH_2] \approx [CH_3NH_3^+]$. So the solution has buffering properties in the proximity of point A. When HBr is added, the H_3O^+ is consumed:

$H_3O^+ + CH_3NH_2 \rightarrow CH_3NH_3^+ + H_2O$

Since the H_3O^+ is removed from the solution (neutralised), the pH does not change significantly.

(a) Explain why the pH does not change significantly between the addition of 5 to 15 mL of HBr (around point **A** on the curve).

Include any relevant equation(s) in your answer.

2011 Exam Q4

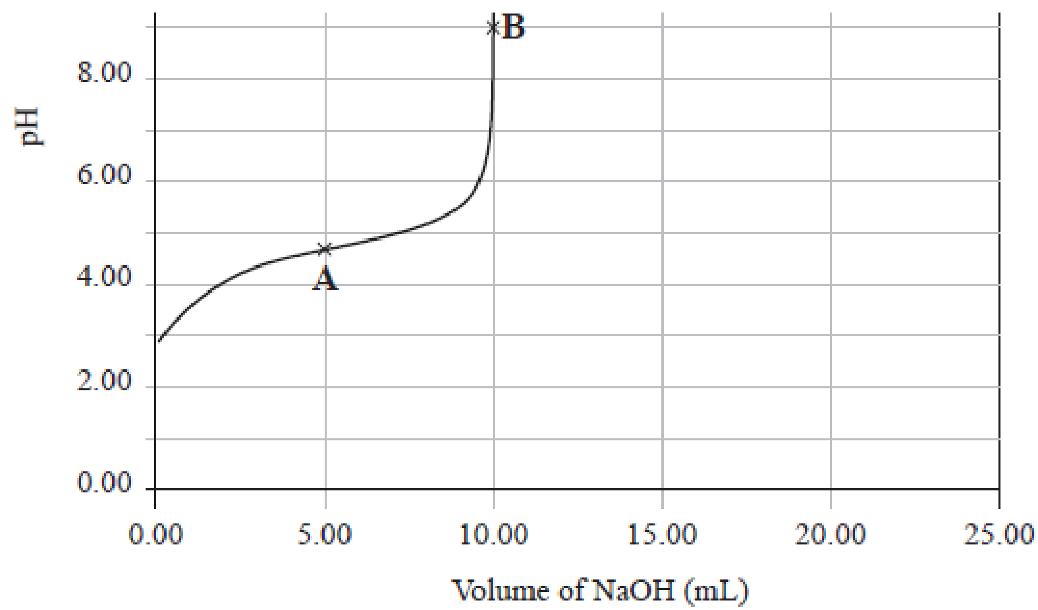
QUESTION FOUR

(a) With reference to the point marked **A** on the graph, discuss:

- the species present, and their relative concentrations
- an estimate of the pK_a value for ethanoic acid
- the effect of adding small amounts of strong acid or strong base to the solution.

Include relevant equations in your answer.

No calculations are necessary.



2011 Exam Q4

ONE of:

- Recognises that at point A there is a buffer solution.
- States that equimolar amounts of acid / base conjugate are present at A.
- States that pH will not change when small amounts of acid or base are added.
- Correct pK_a / K_a

Describes how a buffer works (for when both acid AND base are added) by:
EITHER

- Giving equations for the specific buffer

OR

- Writing about how a buffer works in general terms

OR

- Links that due to equimolar HEt and Et^- thus pK_a = pH

Shows recognition of equimolar HEt and Et^- thus pK_a = pH and discusses how the buffer solution works and links to equations.

A
At point A, there is an equi-molar mixture of HEt and Et^-. On addition of OH^- ions, the acid part of the buffer neutralises the OH^- ions, by donating a proton. The acid reacts with the base:
$HEt + OH^- \rightarrow Et^- + H_2O$

On addition of H_3O^+, the ethanoate will accept a proton from the hydronium ion:
$Et^- + H_3O^+ \rightarrow HEt + H_2O$

Candidate may discuss equilibrium shift.
pK_a = pH = 4.76 (accept 4.5 – 4.9)

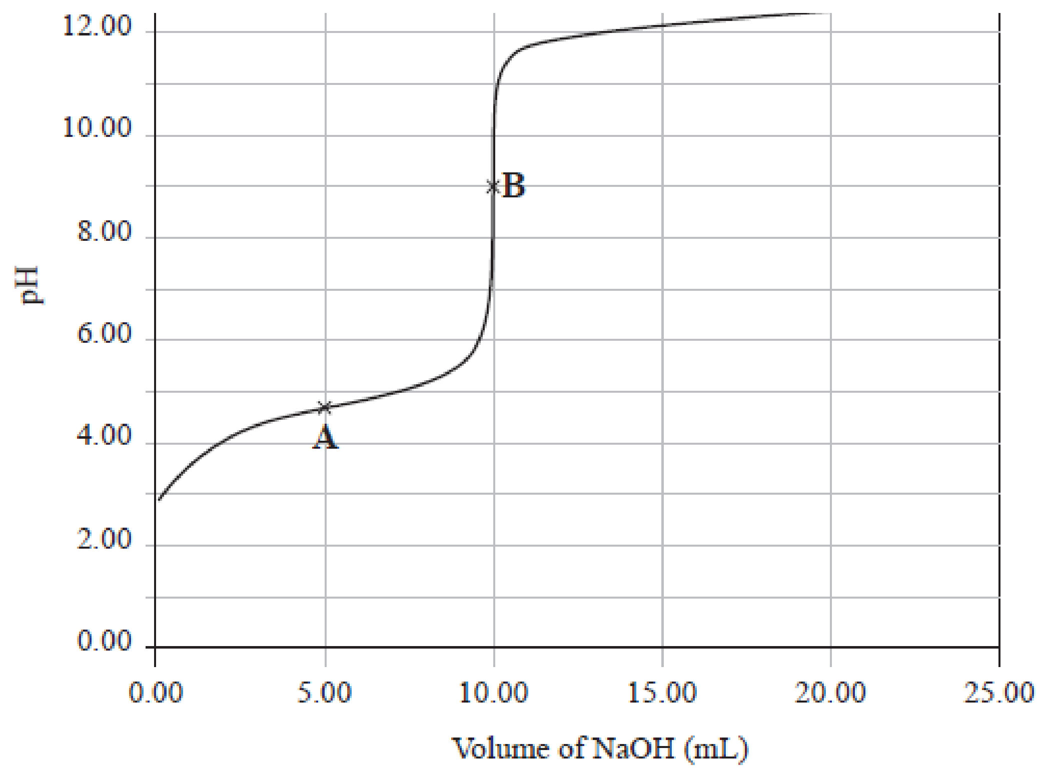
2011 Exam Q4

QUESTION FOUR

(b) With reference to the point marked **B** on the graph, discuss the species present, and their effect on the pH at the equivalence point.

Include relevant equations in your answer.

No calculations are necessary.



2011 Exam Q4

ONE of:

- Recognises that all the HEt has been used up at B.
- That the pH of equivalence point is greater than 7. (must have clearly indicated that point B is the equivalence point)

- Recognises that none of the original HEt remains as ithas all reacted with NaOH

OR

- That the pH of equivalence point is greater than 7 with a valid reason.

Uses two equations to explain why the pH is above 7. (One equation may be implied in the candidate's written answer.)

B

At the equivalence point all the HEt has been neutralised by NaOH.

$$HEt + NaOH \rightarrow EtNa + H_2O$$

The Et^- reacts further to a small extent with water.

$$Et^- + H_2O \rightleftharpoons HEt + OH^-$$

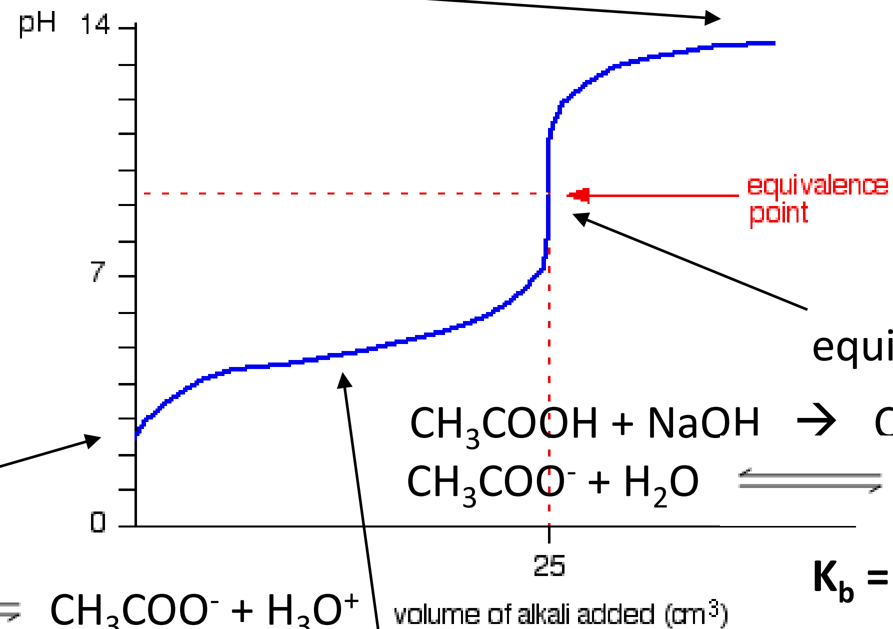
Thus the pH of the equivalence point is above 7 due to presence of OH^-.

Titration curve pH calculations

final pH

$pH = -log[H_3O^+]$

need to recalculate concentrations as volume has changed



need to recalculate concentrations as volume has changed

equivalence point

$CH_3COOH + NaOH \rightarrow CH_3COO^- + H_2O + Na^+$

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

$K_b = [OH^-]^2/[A^-]$

only CH_3COO^-, H_2O and Na^+ present in solution

initial pH

$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$

$K_a = [H_3O^+]^2/[CH_3COOOH]$

½ equivalence

$CH_3COOH + NaOH \rightarrow CH_3COO^- + H_2O + Na^+$

$pH = pK_a$ as $[CH_3COOH] = [CH_3COO^-]$

2013 Sample Exam Q3 b i

QUESTION THREE

20.00 mL of 0.125 mol L^{-1} ethanoic acid is titrated with 0.125 mol L^{-1} sodium hydroxide solution.
The equation for this reaction is:

$$CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$$

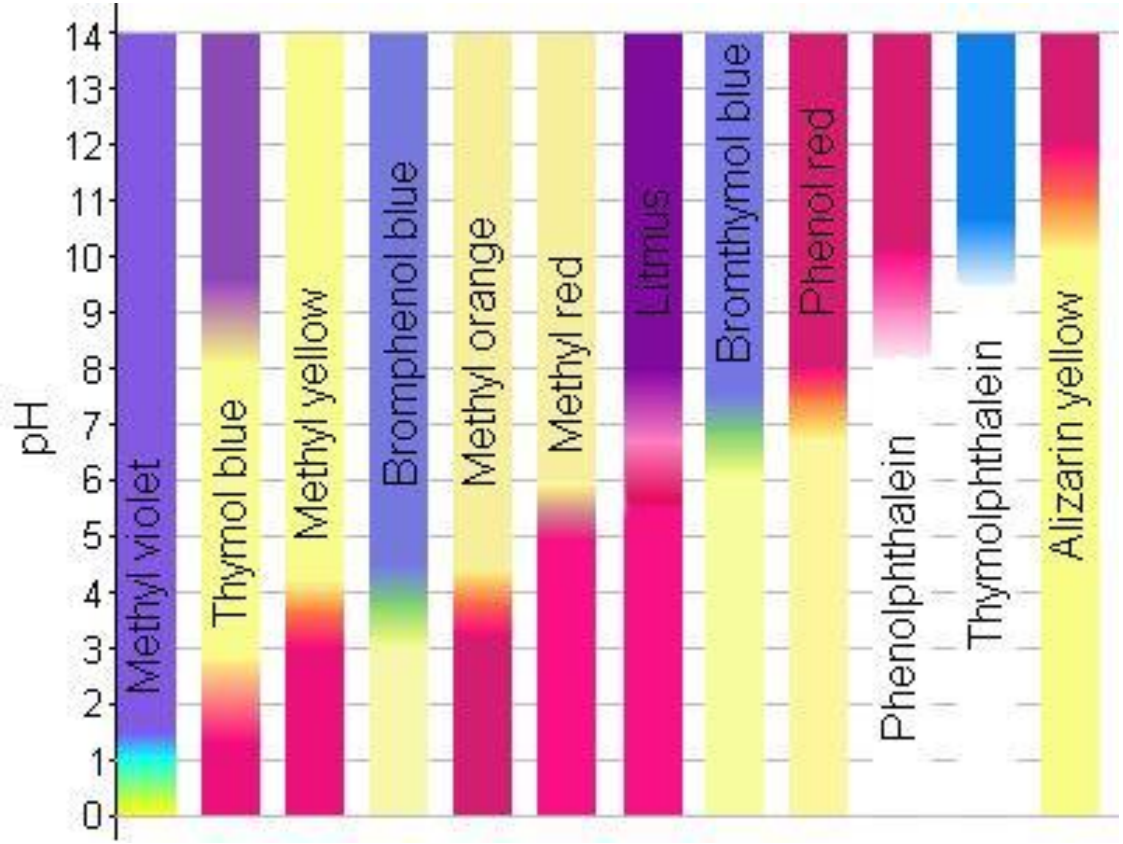
The titration curve for the reaction is given below and the buffer region is marked on the graph.

(b) (i) Show that the pH at the equivalence point for this titration is 8.78.

$$pK_a(CH_3COOH) = 4.76$$

Choosing an indicator

What are we looking for when we choose an indicator for a titration?



Indicators change colour when the pH of the solution = their pK_a

Choosing an indicator

QUESTION THREE

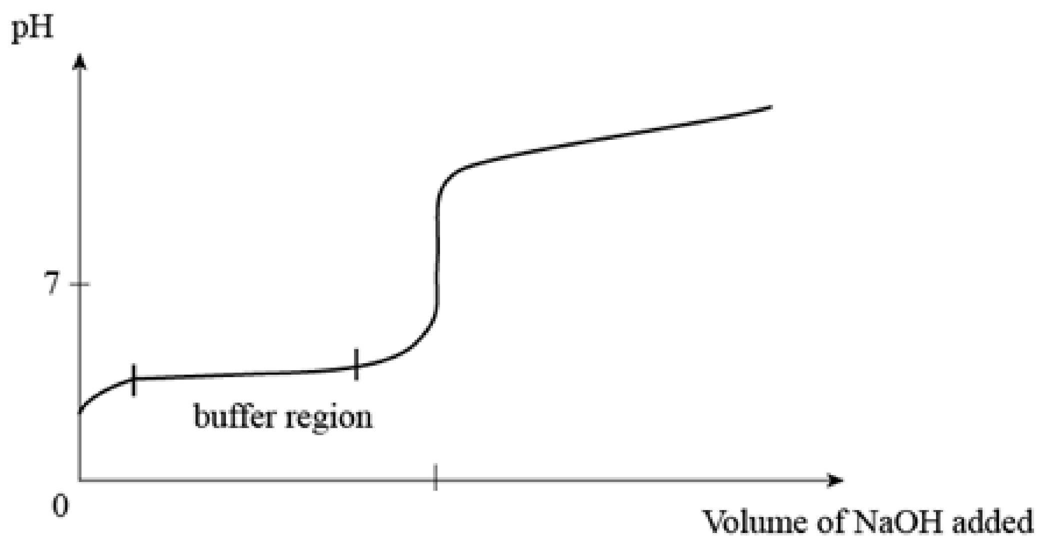
20.00 mL of 0.125 mol L^{-1} etha
The equation for this reaction i

CH_3COO

The titration curve for the react

(ii) Explain why methyl orange is not a suitable indicator for this titration and why phenolphthalein is a suitable indicator for this titration.

Indicator	**pK_a**
Methyl orange	3.70
Phenolphthalein	9.60



pH calculations

Sometime you will be asked to calculate the pH at any point on the titration curve. To do this we need to keep in mind the number of moles of each species, volumes and concentrations.

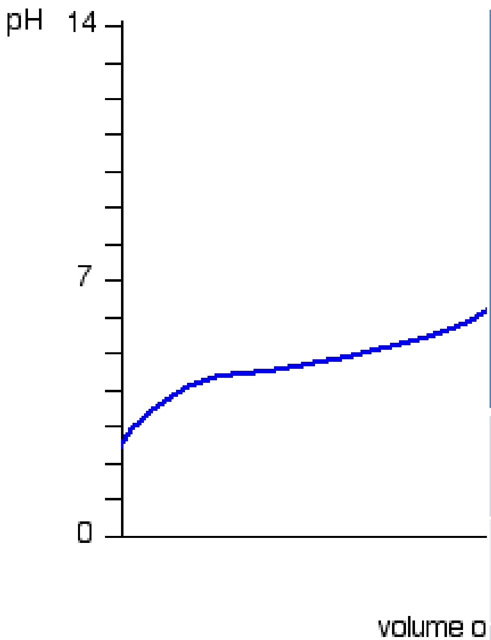
The best way to do this is to use an ICE table.

	Species in flask	Species added to flask	Species formed from reaction
I (initial)			
C (change)			
E (end)			

pH calculations

20 mL of 0.095 mol.L^{-1} CH_3COOH is titrated with 0.110 mol.L^{-1} NaOH.
$K_a(CH_3COOH) = 1.8 \times 10^{-5}$

What is the pH when 10 mL of NaOH has been added to the solution?



	Species in flask (CH_3COOH)	Species added to flask (NaOH)	Species formed from reaction (CH_3COO^-)
I (initial)	**n = 0.0019 mol**	**n = 0 mol**	**n = 0 mol**
C (change)	**n = I – n(NaOH)**	**n = 0.0011 mol**	**n = n(NaOH)**
E (end)	**n = 0.0008 mol**	**n = 0 mol**	**n = 0.0011 mol**

c = 0.02666 mol.L^{-1}

c = 0.03666 mol.L^{-1}

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

pH calculations

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

(ii) Calculate the pH of the titration mixture after 5.00 mL of NaOH has been added.

	CH_3COOH	NaOH	CH_3COO^-
I (initial)	n = 0.001792 mol	n = 0 mol	n = 0 mol
C (change)	n = I – n(NaOH)	n = 0.0005 mol	n = n(NaOH)
E (end)	n = 0.001292 mol	n = 0 mol	n = 0.0005 mol

c = 0.05168 mol.L^{-1} c = 0.02 mol.L^{-1}

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