Acid / Base Worked Example

3 i Calculate the pH of a 0.075 moldm^{-3} solution of sulphuric acid, assuming it to be fully ionised according to the equation

$$H_2SO_4(aq) \rightleftharpoons 2H^+(aq) + SO_4^{2-}(aq)$$

If fully ionized, $[H^+]$ is twice the original acid concentration.

i.e. $[H^+] = 2 \times 0.075$ moldm^{-3}

$= 0.15$ moldm^{-3}

$pH = -\log_{10}[H^+]$

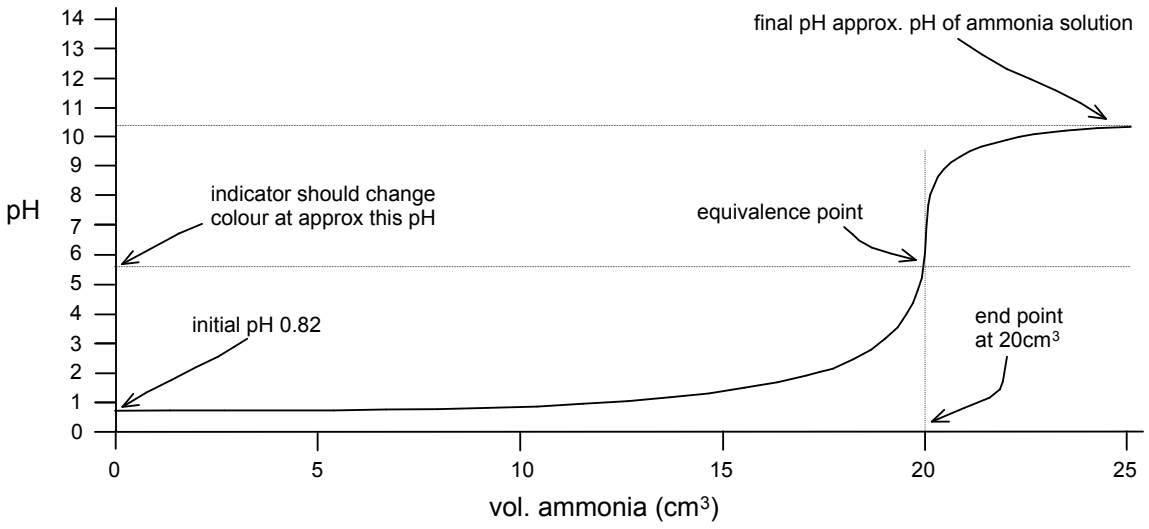
$= -\log_{10}(0.15)$

$= 0.82$

ii A total of 25.0cm^3 of ammonia solution was added in small portions from a burette to 10.0cm^3 of 0.075moldm^{-3} sulphuric acid.

The pH of the solution was followed as the ammonia was added.

Sketch a graph showing how the pH changed, assuming that 20.0cm^3 of the ammonia solution was sufficient to neutralise the fully ionised sulphuric acid.



PTO

iii Calculate the concentration in moldm^{-3} of the ammonia solution.

Equation	$2NH_3$ + H_2SO_4 → $2NH_4^+$ + SO_4^{2-}	
Ratio	2 : 1	
Moles	$2 \times 7.5 \times 10^{-4}$ = 1.5 x 10-3 moles	$\frac{0.075 \times 10}{1000}$ = 7.5×10^{-4}

moles H_2SO_4 = concentration x volume / 1000

moles NH_3 is twice that of H_2SO_4

This is in 20cm^3 at end point

1.5 x 10^{-3} moles	:	20cm^3	/ 20 x 1000
0.075 moles	:	1000cm^3	

Therefore $[NH_3]$ = 0.075 moldm^{-3}

iv Not all indicators are suitable for this titration. Explain why some indicators cannot be used.

Would want an indicator which changes colour just below pH7.
This is because the equivalence point is at about pH5.5.

A suitable indicator would be methyl orange