A buffer solution if made when 1.50 g of sodium hydroxide is added to 1.00 dm^3 of a 0.150 moldm^{-3} solution of a weak acid, HA.

The acid dissociation constant, Ka = 1.79 x 10-5 moldm^{-3}

1dm^3 is 1000cm^3

c) Calculate the pH of this buffer solution.

① Write an expression for Ka:

$$Ka = \frac{[H^+][A^-]}{[HA]}$$ products over reactants

② Find the moles of acid and base used:

moles of NaOH $= \frac{1.50}{40} = 0.0375$

moles of HA $= 0.150$

moles $= \frac{\text{mass}}{\text{Mr}}$

③ Use this to calculate the moles reacted:

Acid in excess

$\Rightarrow$ 0.0375 moles react.

④ Calculate the moles of each species at equilibrium:

	$HA \rightleftharpoons$	H^+ +	A^-
initial moles	0.150	0	0
equilibrium moles	0.1125		0.0375

conc. = $\frac{\text{moles} \times 1000}{\text{vol.}}$

⑤ Find the concentrations at equilibrium:

$[HA] = \frac{0.1125 \times \cancel{1000}}{\cancel{1000}} = 0.1125\ moldm^{-3}$

$[A^-] = \frac{0.0375 \times \cancel{1000}}{\cancel{1000}} = 0.0375\ moldm^{-3}$

⑥ Sub these value into the Ka expression to find [H+] ions:

$$K_a = \frac{[H^+][A^-]}{[HA]} \Rightarrow [H^+] = \frac{K_a[HA]}{[A^-]}$$

$$\Rightarrow [H^+] = \frac{1.79 \times 10^{-5} \times 0.1125}{0.0375}$$

$$= 5.37 \times 10^{-5}\ mol\,dm^{-3}$$

⑦ Calculate the pH of the buffer solution:

$$pH = -\log_{10}(5.37 \times 10^{-5})$$

$$= 4.2700...$$

$$= \underline{4.27}$$

the pH of buffers is usually around 4.

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