Marks 4

- Buffer 1 is a solution containing 0.08 M NH_4Cl and 0.12 M NH_3. Buffer 2 is a solution containing 0.15 M NH_4Cl and 0.05 M NH_3. The acid dissociation constant of the ammonium ion is 5.50×10^{-10}. What are the pH values of each of the buffer solutions?

By definition, $pK_a = -\log_{10}K_a$ so:

$$pK_a = -\log_{10}(5.50 \times 10^{-10}) = 9.26$$

Using the Henderson-Hasselbalch equation,

$$pH = pK_a + \log\frac{[\text{base}]}{[\text{acid}]} = 9.290 + \log\frac{[NH_3]}{[NH_3Cl]}$$

For buffer 1:

$$pH = 9.26 + \log\frac{0.12}{0.08} = 9.44$$

For buffer 2:

$$pH = 9.26 + \log\frac{0.05}{0.15} = 8.78$$

Buffer 1 pH = **9.44**	Buffer 2 pH = **8.78**

Which buffer is better able to maintain a steady pH on the addition of small amounts of both a strong acid and strong base? Explain.

Buffer 1 is better able to maintain a steady pH because its pH is closer to the pK_a of NH_4^+. This is because it has relatively high concentrations of both NH_4^+ and NH_3 which can react with any added OH^- or H^+ respectively.