- HA dissociates to release H^+ ions
- The position of equilibrium shifts to the right
- H^+ concentration has returned to normal

Everyday applications of buffers include **reactive dyes on fabrics, maintaining blood pH.**

<u>Finding pH from buffers</u>

As we have a weak acid in the solution of a buffer, we can use the acidity constant equation:

$$Ka = \frac{[H^+] \times [A^-]}{[HA]}$$

If we think about our assumptions, A^- will come purely from the salt, so can be represented by the salt concentration.

HA will remain completely unchanged and can therefore be represented by the concentration of acid on addition. This creates the equation:

$$Ka = [H^+] \times \frac{[salt]}{[acid]}$$

The **value of H^+** is affected by the value of **Ka**, and the ratio of **salt : acid.**

This also means pH will not be influenced by dilution, as diluting the solution changes the concentration of salt and acid equally.

With a Ka value of 1.7×10^{-5}, and $0.1dm^{-3}$ ethanoic acid with $0.2dm^{-3}$ sodium ethanoate, we can calculate pH:

$$1.7 \times 10^{-5} = [H^+] \times \frac{0.2}{0.1}$$

The equation is rearranged to make H^+ the subject by **switching salt and acid** around, and **switching H^+ and Ka:**

$$H^+ = 1.7 \times 10^{-5} \times \frac{0.1}{0.2} = 8.5 \times 10^{-6}$$

We then input the H^+ concentration into the pH equation:

$$-\log (8.5 \times 10^{-6}) = \boldsymbol{pH\ 5.07}$$

Explain hydrogen bonding in water and the unusual properties that arise from it

There are 3 conditions which must be met for hydrogen bonding to occur:

- A hydrogen must be covalently bonded to a highly electronegative element (O, N, F) to produce a **large dipole**
- The **small hydrogen atom** must be close to an O, N or F on another nearby molecule
- A **lone pair of electrons** on the O, N or F for hydrogen to line up with