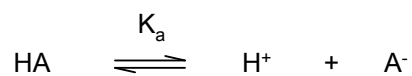
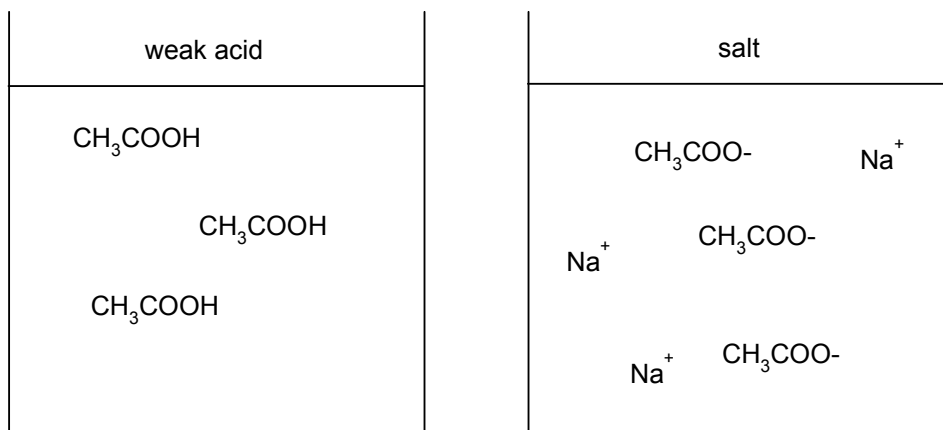


Buffer Solutions Examples

- 1 a) What is the pH of a buffer solution containing 0.1 mol dm^{-3} ethanoic acid and 0.1 mol dm^{-3} sodium ethanoate? (pK_a ethanoic acid = 4.75)



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

$\xrightarrow{\times [\text{HA}]}$ $K_a [\text{HA}] = [\text{H}^+][\text{A}^-]$

$\xrightarrow{\text{swap sides}}$ $[\text{H}^+][\text{A}^-] = K_a [\text{HA}]$

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

$$\text{pH} = \text{pK}_a + \text{p} \frac{[\text{HA}]}{[\text{A}^-]}$$

$$= 4.75 - \log(0.1 / 0.1)$$

$$= \underline{\underline{4.75}}$$

note that $\text{pH} = \text{pK}_a$ when the concentration of weak acid and salt is the same

b) What is its pH after the addition of 10cm³ of 1.0 moldm⁻³ HCl(aq) to 1dm³ of the solution?

How many moles of HCl are added?

$$\begin{aligned} \text{moles} &= \frac{\text{concentration} \times \text{volume}}{1000} \\ &= \frac{1.0 \times 10}{1000} \\ &= 0.01 \text{ moles HCl added} \\ &= 0.01 \text{ moles H}^+ \text{ added} \end{aligned}$$

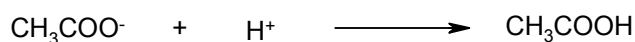
[CH₃COOH] = 0.1 moldm⁻³
[CH₃COO⁻] = 0.1 moldm⁻³
and there is 1dm³

Equation	CH ₃ COOH	⇌	CH ₃ COO ⁻	+	H ⁺
Ratio	1	:	1	:	1
Initial Moles	$\frac{0.1 \times 1000}{1000}$ = 0.1 moles		$\frac{0.1 \times 1000}{1000}$ = 0.1 moles		
Final Moles	0.1 + 0.01 = 0.11 moles		0.1 - 0.01 = 0.09 moles		

Amount of acid is increased by the amount of H⁺ added

Amount of base is reduced by the amount of H⁺ added

This is because the base will have reacted with the H⁺ from HCl to form more ethanoic acid:



i.e. the reverse reaction in the equilibrium is favoured.

$$[\text{CH}_3\text{COOH}] = \frac{0.11}{1010}$$

$$[\text{CH}_3\text{COO}^-] = \frac{0.09}{1010}$$

The volume is now 1dm³ for the original volume plus the 10cm³ for the HCl added

$$\text{pH} = \text{pKa} + \text{p} \frac{[\text{HA}]}{[\text{A}^-]}$$

$$= 4.75 - \log \frac{(0.11 / 1010)}{(0.09 / 1010)}$$

$$= \underline{\underline{4.66}}$$

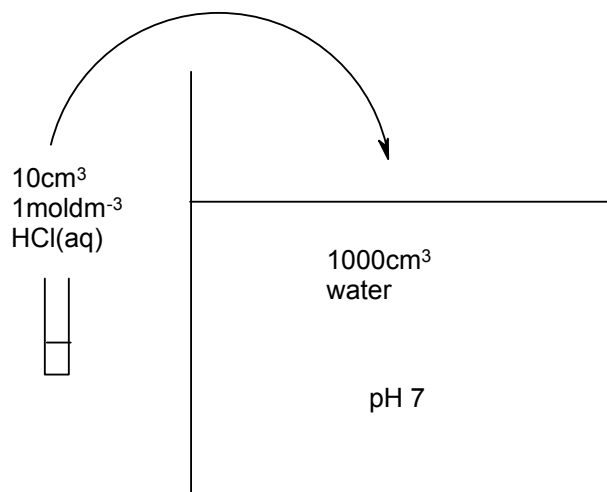
but in fact the volumes cancel in this equation

The pH has gone down from 4.75 to 4.66.
The buffer has resisted change in pH on addition of a small amount of acid.

- c) If 10cm^3 of 1mol dm^{-3} HCl had been added to 1dm^3 of water, what would the change in pH be? (assume pH of pure water at this temperature is 7)

Moles of H^+ added:

$$\begin{aligned} \text{moles} &= \frac{\text{concentration} \times \text{volume}}{1000} \\ &= \frac{1.0 \times 10}{1000} \\ &= 0.01 \text{ moles HCl added} \\ &= 0.01 \text{ moles H}^+ \text{ added} \end{aligned}$$



This is in 1010cm^3 of water (1dm^3 plus the 10cm^3 of acid solution).

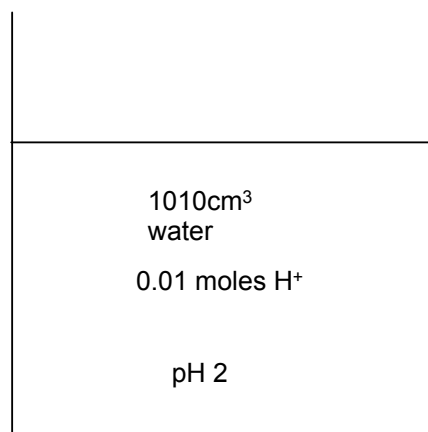
Therefore the concentration of H^+ is:

$$\begin{aligned} \text{concentration} &= \frac{\text{moles}}{(\text{volume} / 1000)} \\ [\text{H}^+] &= \frac{0.01}{(1010 / 1000)} \\ &= 9.9 \times 10^{-3} \text{ mol dm}^{-3} \end{aligned}$$

$$\text{pH} = -\log [\text{H}^+]$$

$$= -\log 9.9 \times 10^{-3}$$

$$= \underline{2.0}$$



Therefore the pH has reduced from 7 to 2. This is a change of 5 on the pH scale. Compare this with the change of pH in the buffer solution. The pH reduced from 4.75 to 4.66. This is a change of only 0.09 on the pH scale.