Name:

pH and Indicators	Objectives
18. pH and Indicators	-define pH -describe the use of the pH scale as a measure of the degree of acidity/alkalinity -discuss the limitations of the pH scale **-explain self-ionisation of water** -write an expression for Kw -use universal indicator paper or solution to determine pH -calculate the pH of dilute aqueous solutions of strong acids and bases **-distinguish between the terms weak, strong, concentrated and dilute in relation to acids and bases** **-calculate the pH of weak acids and bases (approximate method of** **-calculation to be used – assuming that ionisation does not alter the total concentration of the non-ionised form)** **-define acid-base indicator** **-explain the theory of acid-base indicators** -justify the selection of an indicator for acid base titrations

Water only conducts electricity when it contains ions dissolved in it. However, even pure water conducts a very small current, so there must be some ions present. The ions exist because water *self-ionises:*

$$H_2O \rightleftharpoons H^+ + OH^-$$

The equilibrium above lies *far* to the right i.e. only a very small number of H_2O molecules self-ionise.

As the number of H_2O molecules that self-ionise is so small, we can say that the concentration of H_2O stays constant.

Look at the K_c expression for the self-ionisation of water:

$$K_c = \frac{[H^+][OH^-]}{[H_2O]}$$

K_c is a constant. We also assumed that $[H_2O]$ is a constant, so we can bring the $[H_2O]$ up and multiply on the left side.

$$K_c[H_2O] = [H^+][OH^-]$$

K_c (a constant) times $[H_2O]$ (a constant), gives a constant. We call this constant K_w, the *ionic product of water*.

Def^n: The **ionic product of water** is $K_w = [H^+][OH^-]$

For water at 25°C it has been found that K_w is $1\text{x}10^{-14}$. This means

$$K_w = [H^+][OH^-] = 1 \times 10^{-14}$$

To find $[H^+]$, we need to use the fact that $[H^+]=[OH^-]$ (from the 1:1 ratio in the self-ionisation equation).

$$K_w = [H^+][H^+] = 1 \times 10^{-14}$$

$$K_w = [H^+]^2 = 1 \times 10^{-14}$$

$$[H^+] = \sqrt{1 \times 10^{-14}}$$

$$[H^+] = 1 \times 10^{-7}$$

See how this could be asked in HL 2016 Q10 a (iii)

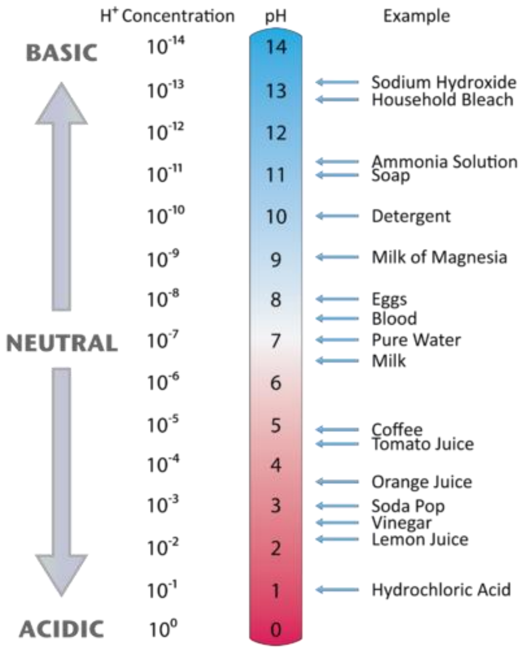
G. Galvin

Calculating pH:

Def^{n}: $\boldsymbol{pH = -log_{10}[H^+]}$

For strong acids:

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



Find the pH of a solution whose H^+ concentration is 7.4×10^{-7} mol/L

Find the pH of a 0.11 mol/L solution of HCl

Find the pH of a 0.5 mol/L solution of H_2SO_4

Find the H^+ concentration of a HCl solution whose pH is 1.9

For strong bases:

$$pOH = -log_{10}[OH^-]$$

and

$$pH = 14 - pOH$$

Find the pH of a solution whose OH^- concentration is 4.6×10^{-9} mol/L

Find the pOH of a 0.35 mol/L solution of NaOH

Find the pH of a 0.75 mol/L solution of $Ca(OH)_2$

Find the OH^- concentration of an KOH solution whose pH is 13.2

For weak acids:

$\boldsymbol{pH = -log_{10}\sqrt{M_{acid} \times K_a}}$ where M_{acid} is the molarity of the acid and K_a is the dissociation constant of the acid (given in question).

Find the pH of a 0.1M solution of methanoic acid given that K_a is 2.1×10^{-4}

Find the pH of an aqueous solution containing 1.48 g of propanoic acid (CH_3CH_2COOH) in 200 cm^3 in solution, given that the dissociation constant K_a for propanoic acid is 1.36×10^{-5}.

Find the molarity of a CH_3COOH solution whose pH is 1.9, given the K_a value for CH_3COOH is 1.78×10^{-5}.

For weak bases:

$pOH = -log_{10}\sqrt{M_{base} \times K_b}$ where M_{base} is the molarity of the base and K_b is the dissociation constant of the base (given in question). Also remember that pH = 14 - pOH

Find the pH of a 0.15M solution of ammonia (NH_3) given that K_b is 1.8×10^{-5}

Find the pH of an aqueous solution containing 1.5 g of the basic compound $C_2H_5NH_2$ in 250 cm^3 in solution, given that the dissociation constant K_b for this base is 4.4×10^{-4}.

Find the molarity of a NH_3 solution whose pH is 12, given the K_b value for CH_3COOH is 1.8×10^{-5}.

Limitations of the pH scale:

1. Only has a range of 0-14.
2. Does not work for concentrated solutions.
3. Only works for aqueous solutions.

Acid/Base Indicators:

Indicators are weak acids/bases that change colour depending on where the equilibrium lies between the undissociated and dissociated forms.

1) *Indicator as a weak acid:*

$$\underset{\text{Colour 1}}{HIn} \rightleftharpoons H^+ + \underset{\text{Colour 2}}{In^-}$$

By Le Chatelier's Principle,
1. Adding H^+ (acid) will favour the reverse reaction, so Colour 1 will be seen.
2. Adding OH^- (base) will react with the H^+ to form water. As this removes H^+ from the system, the forward reaction will be favoured, so Colour 2 will be seen.

2) *Indicator as a weak base:*

$$\underset{\text{Colour 1}}{XOH} \rightleftharpoons \underset{\text{Colour 2}}{X^+} + OH^-$$

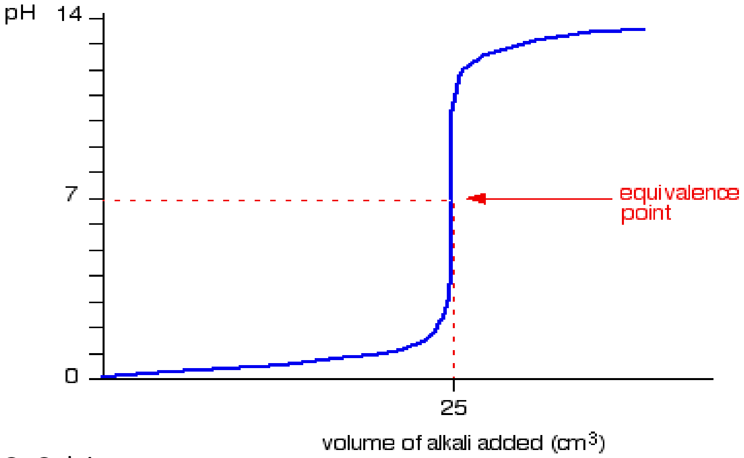
By Le Chatelier's Principle,
1. Adding OH^- (base) will favour the reverse reaction, so Colour 1 will be seen.
2. Adding H^+ (acid) will react with the OH^- to form water. As this removes OH^- from the system, the forward reaction will be favoured, so colour 2 will be seen.

Indicators are chosen based on the pH range in which the colour changes. The indicators and pH ranges we need to know are:

Name of indicator	Approx. pH range	Acid Colour	Base Colour
Methyl Orange	3-5	Red	Yellow
Litmus	5-8	Red	Blue
Phenolphthalein	8-10	Colourless	Pink

Titration Curves:

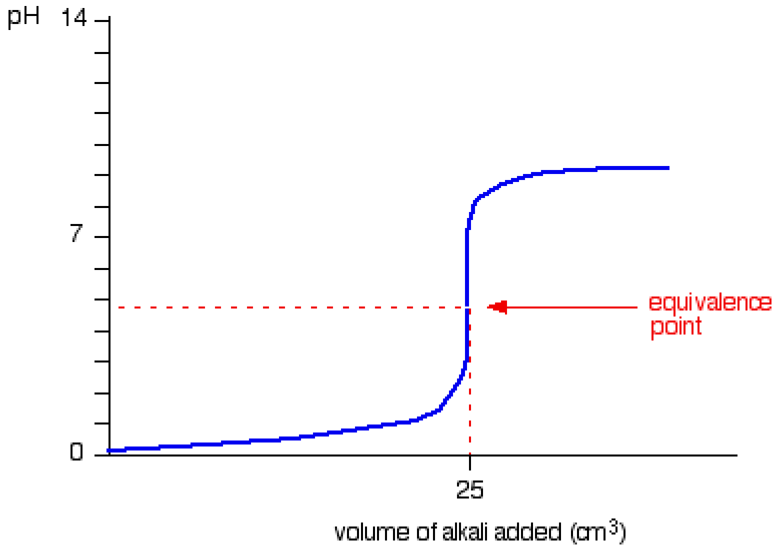
Strong Acid vs. Strong Base



Notes:

1. pH begins near 0.
2. pH finishes near 14.
3. Long vertical section where solution quickly changes from acidic to basic.
4. Midpoint of this vertical section is the "equivalence point".
5. Indicator chosen must completely change colour within the vertical section in order to be suitable (methyl orange & phenolphthalein both suitable here)

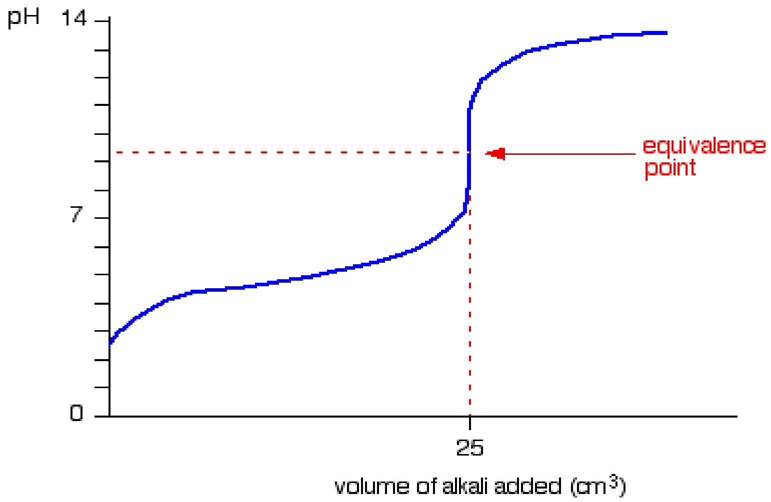
Strong Acid vs. Weak Base



Notes:

1. pH begins near 0.
2. pH finishes much lower than 14.
3. Shorter vertical section where solution quickly changes from acidic to basic.
4. Midpoint of this vertical section is the "equivalence point".
5. Indicator chosen must completely change colour within the vertical section in order to be suitable (methyl orange only suitable here)

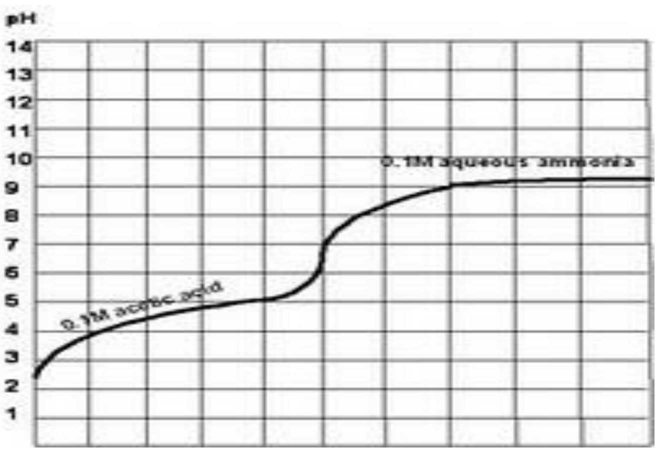
Weak Acid vs. Strong Base



Notes:

1. pH begins much higher than 0.
2. pH finishes near 14.
3. Shorter vertical section where solution quickly changes from acidic to basic.
4. Midpoint of this vertical section is the "equivalence point".
5. Indicator chosen must completely change colour within the vertical section in order to be suitable (phenolpthalein only suitable here)

Weak Acid vs. Weak Base



Notes:

1. pH begins much higher than 0.
2. pH finishes much lower than 14.
3. No vertical section where solution quickly changes from acidic to basic.
4. No obvious "equivalence point".
5. Indicator chosen must completely change colour within the vertical section in order to be suitable (no indicator suitable here)