pH and Indicators

Self Ionisation of Water

It has been found that even pure water will conduct a tiny amount of electricity. If the water was pure then where are the ions coming from?

The answer is that water itself dissociates slightly to from ions.

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

$$2H_2O \rightleftharpoons H_3O^+ + OH^-$$

Ionic Product of Water

$K_w = K_c\,[H_2O] = [H^+][OH^-]$

At 25^0C $K_w = 1.0 \times 10^{-14}$

So $[H^+][OH^-] = 1.0 \times 10^{-14}$ (also $[H^+] = [OH^-]$ in pure water)

Then $[H^+]^2 = 1.0 \times 10^{-14}$

$[H^+] = 1.0 \times 10^{-7}$

pH Scale

$pH = -\log_{10}[H^+]$ where [ ] represents concentration in mol/L

The pH of a solution is the negative logarithm to the base 10 of the hydrogen ion concentration measured in moles per litre.

1	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}	10^{-9}	10^{-10}	10^{-11}	10^{-12}	10^{-13}	10^{-14}

Increasing acidity $[H+] > 10^{-7}$

Increasing alkalinity $[H+] < 10^{-7}$

Calculating pH

What is the pH of a solution which has 10^{-3} moles of H^+ ions per litre?

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

$pH = - \log_{10}[10^{-3}]$

$pH = - (-3)$

$pH = 3$

Calculate the pH of 0.02 moles of HCl

$$HCl \longrightarrow H^+ + Cl^-$$

HCl: H^+

1:1

0.02 moles : 0.02 moles

$-\log_{10}[H^+] = -\log_{10}[0.02] = 1.7$

pH = 1.7

Calculate the pH of 0.02 moles of H_2SO_4

$$H_2SO_4 \longrightarrow 2H^+ + SO_4^{2-}$$

$$H_2SO_4 : 2H^+$$

$$1:2$$

$$0.02 \text{ moles} : 0.04 \text{ moles}$$

$$-\log_{10}[H^+] = -\log_{10}[0.04] = 1.4$$

$$pH = 1.4$$

Strengths of Acids and Bases

A strong acid is a good proton donor.

A weak acid is a poor proton donor.

Strong acid = weak conjugate base.

A strong base is a good proton accepter.

A weak base is a poor proton accepter.

Strong base = weak conjugate acid.

Acid dissociation constant is $K_a = \frac{[H^+][A^-]}{[HA]}$

The larger the value of K_a, the higher the concentration of H_3O^+ and the stronger the acid

pH and pOH

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

$pOH + pH = 14$

So is a solution has a pH of 1.5 then the pOH of that solution is 12.5

Calculate the pH of a solution containing 4g of sodium hydroxide (NaOH)

$NaOH \longrightarrow Na^+ + OH^-$

NaCl: H^+

1:1

0.1 moles : 0.1 moles

$-\log_{10}[OH^-] = -\log_{10}[0.1] = 1$

pOH = 1

pH = 14 – pOH

pH = 14 – 1 = 13

RMM of NaOH = 23 + 16 + 1 = 40g

$$moles = \frac{mass}{RMM} = \frac{4}{40} = 0.1$$

Calculating the pH of weak acids and bases.

Because weak acids and bases do not dissociate fully we need to work out the K_a/K_b of the substance in order to calculate its pH.

Weak acid: $[H^+] = \sqrt{K_a \ x \ M_{acid}}$

Weak base: $[OH^-] = \sqrt{K_b \ x \ M_{base}}$

Concentration of acid/base in moles per litre

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

$[H^+] = \sqrt{K_a \; x \; M_{acid}}$

$= \sqrt{2.1 \times 10^{-4} \; x 0.1}$

$= \sqrt{2.1 \; x 10^{-5}}$

$= 0.0046$

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

$= -\log_{10+}[0.0046]$

$= 2.34$

$pH = 2.34$

Limitations of the pH scale

1. The scale is limited to 0-14 range although pH values outside this range are possible (in theory).

2. pH scale only works for aqueous solutions, not all reactions occur in water.

3. pH scale does not work for highly concentrated solutions (greater than 1M).

Indicators

An acid-base indicator is a substance that changes colour according to the pH of the solution in which it is placed.

The range of an indicator is the pH interval over which there is a clear change in colour for that indicator.

Indicator	pH Range	Acid colour	Base colour
Methyl orange	3-5	red	yellow
Litmus	5-8	red	blue
Phenolphthalein	8-10	colourless	pink

Indicators

Indicators are weak acids or bases that dissociate slightly.

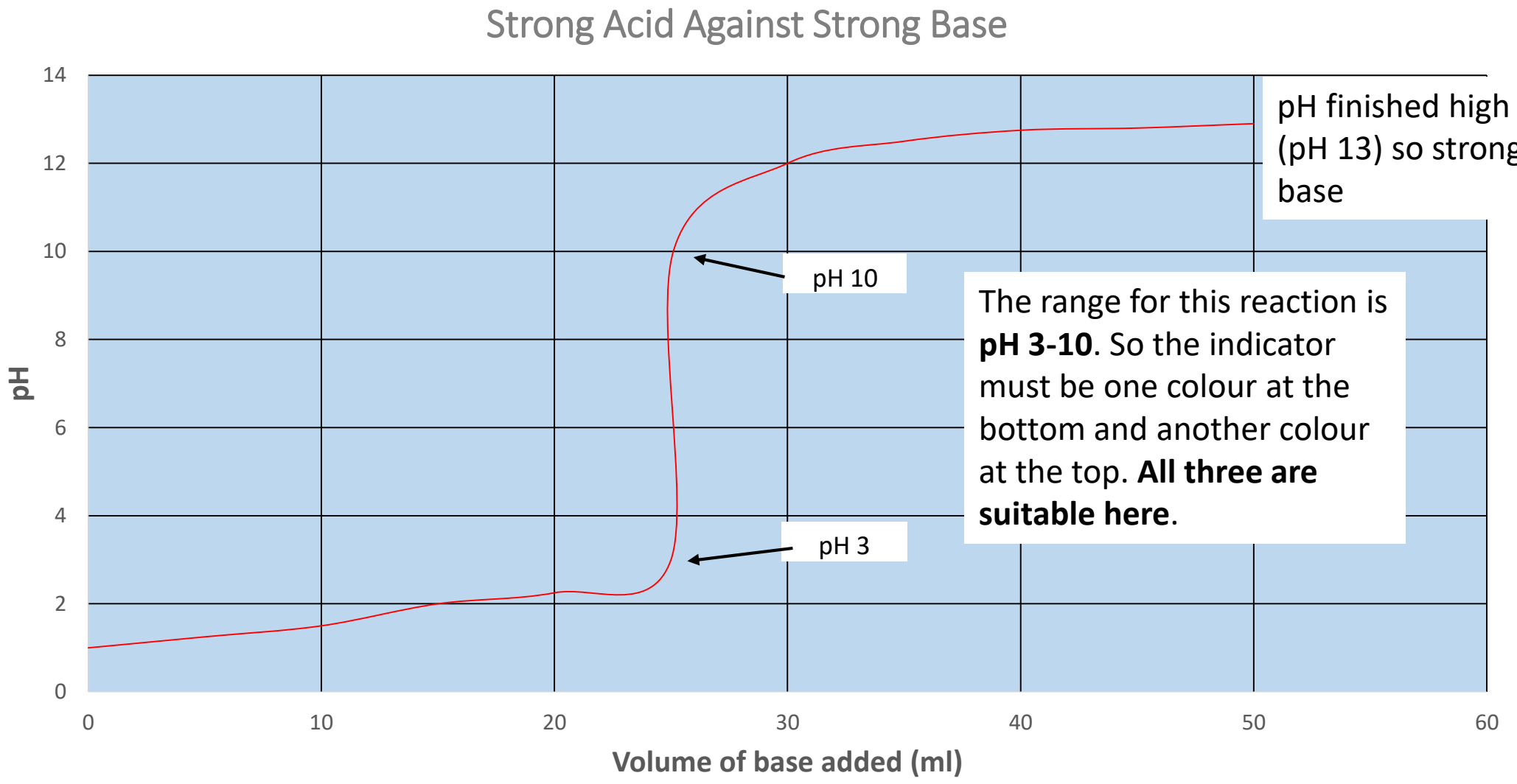
Due to this, only a few drops are used in titrations so that the pH of the indicator doesn’t affect the result.

Universal indicator is a mixture of indicators that gives colour change over the entire pH range.

pH graphs

pH graphs are used to show the change in pH throughout a titration.

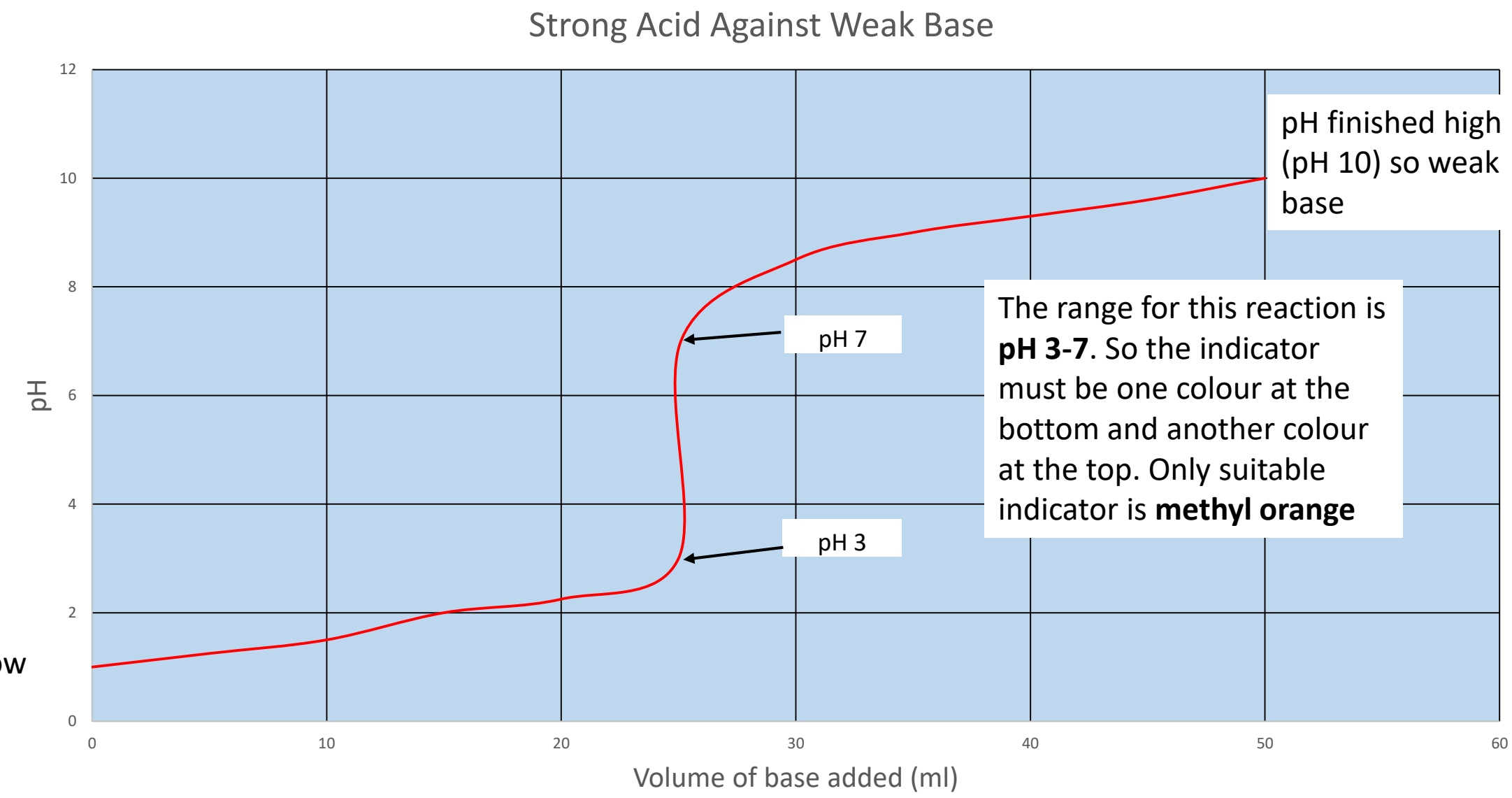
You must be able to draw and interpret these graphs.



pH finished high (pH 13) so strong base

The range for this reaction is **pH 3-10**. So the indicator must be one colour at the bottom and another colour at the top. **All three are suitable here**.

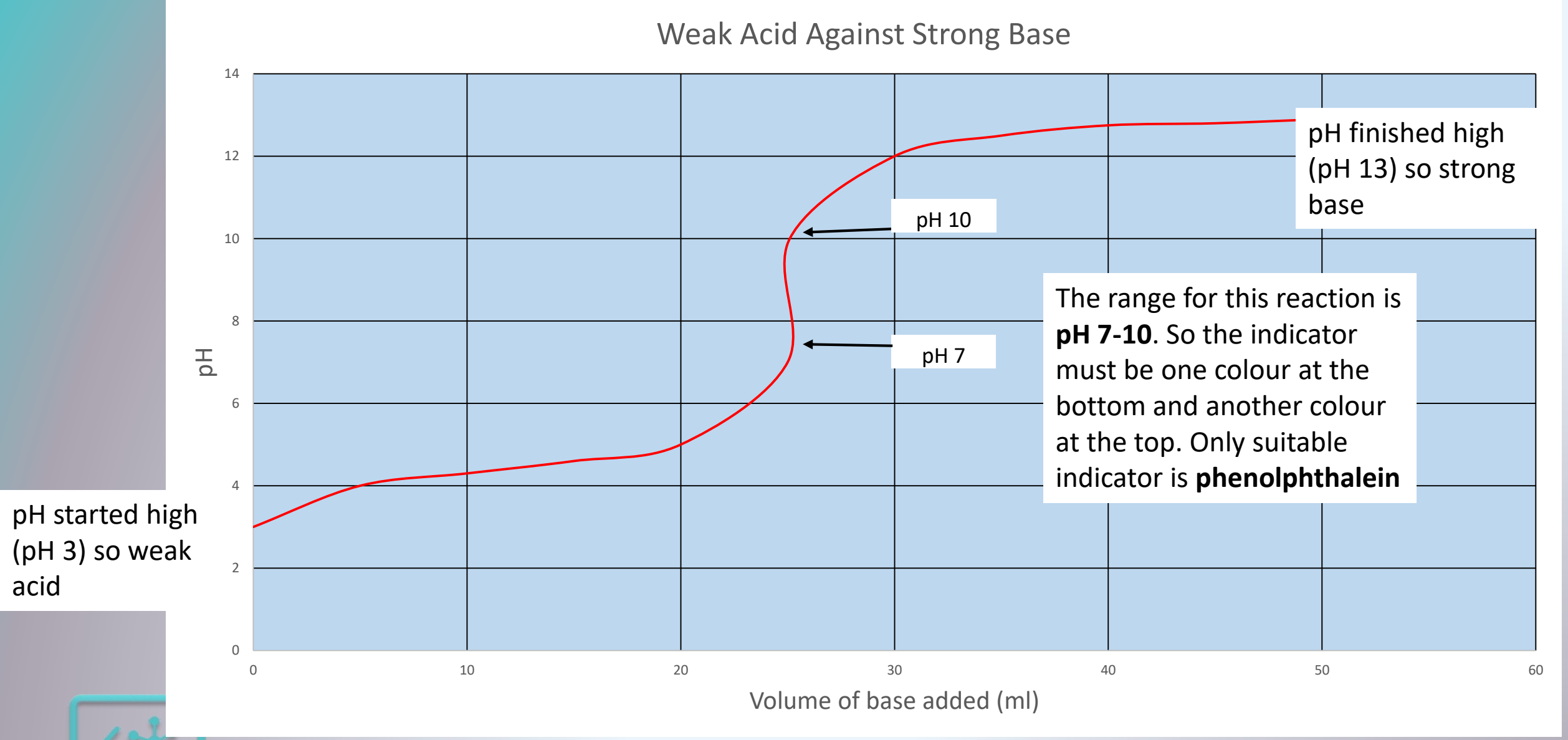
pH started low (pH 1) so strong acid

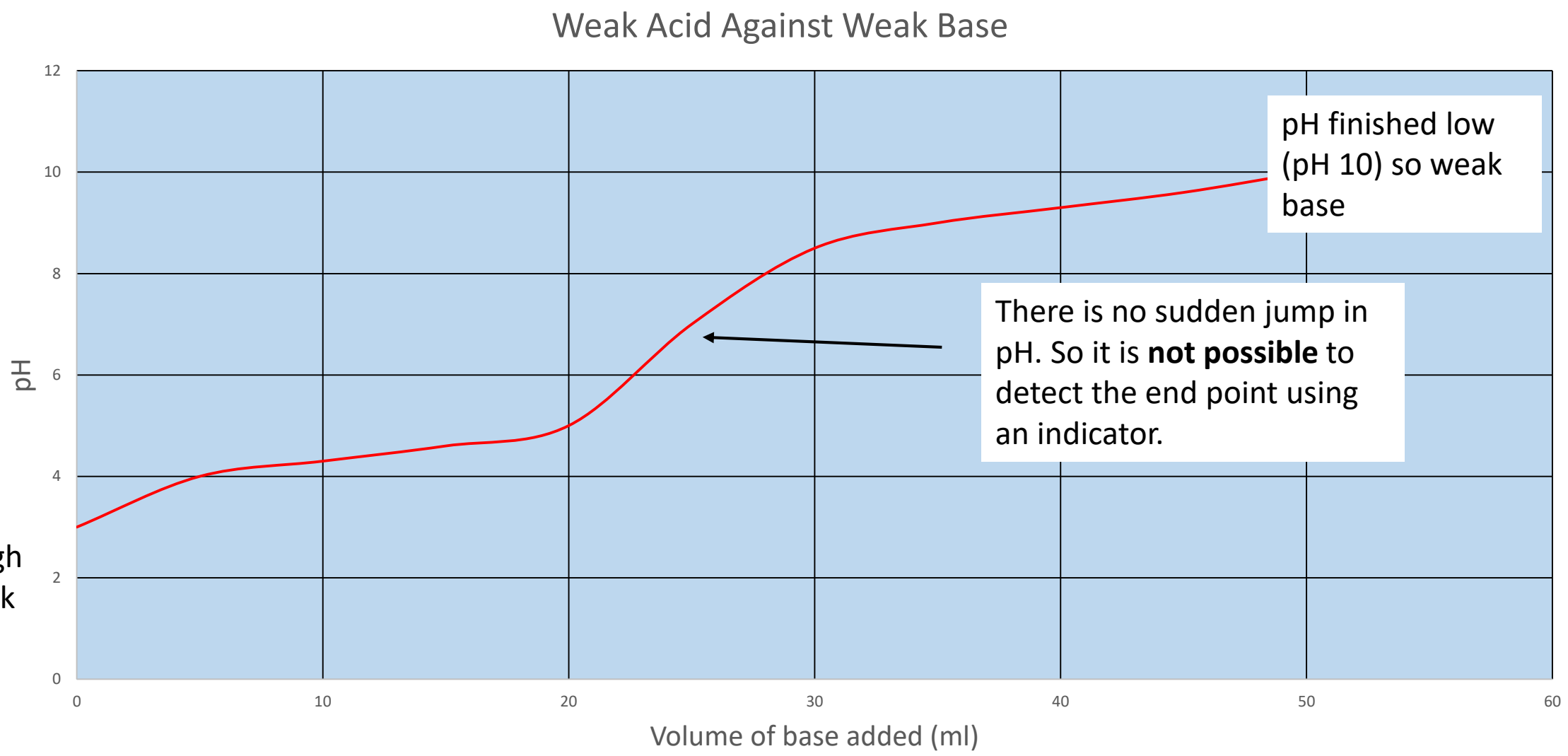


pH finished high (pH 10) so weak base

The range for this reaction is **pH 3-7**. So the indicator must be one colour at the bottom and another colour at the top. Only suitable indicator is **methyl orange**

pH started low (pH 1) so strong acid





pH started high (pH 3) so weak acid

classroomchemistry