Example 2. Observe the changes in pH brought by changes in concentration

Concentration of HCl	***pH***
1	*0*
$1x10^{-1}$	*1*
$1x10^{-2}$	*2*
$1x10^{-3}$	*3*

Strong acid has a weak conjugate base. If an acid has a strong tendency to donate protons, its conjugate base will have a poor tendency to accept protons.
Similarly, weak acids have strong conjugate bases.

Dilution of Strong and Weak Acids

- Strong acid → fully dissociated
  As diluted, the hydrogen ions concentration falls in line with the dilution factor.
  For each dilution of 10x, the pH increases by 1 unit.
- Weak acids → partially dissociated
  As diluted, some undissociated acid molecules split up
  pH does not increase as fast as it does with the strong acid.
  For each dilution of 10x, the pH increases by 0.5 unit,
  For each dilution of 100x, the pH increases by 1

. Dilution factor	Concentration of acid /moldm^{-3}	pH of strong acid (monobasic)	pH of weak acid (monobasic)
0	0.1	1	2.88
10x	0.01	2	3.38
100x	0.001	3	3.88
1000x	0.0001	4	4.38

Acid Dissociation constant K_a

A strong acid is one that dissociated completely. The greater the dissociation the stronger the acid. The amount of dissociation is therefore an indication of the strength of an acid and is measured using the dissociation constant. Since dissociation is and equilibrium process, the equilibrium expression for the dissociation of acid, HA, is:

$$HA(aq) \quad + \quad H_2O(l) \quad \rightleftharpoons \quad H_3O^+(aq) \quad + \quad A^-(aq)$$

The equilibrium expression K_c for the dissociation of a weak acid

$$K_c = \frac{[H_3O^+][A^-]}{[HA]\,[H_2O]}$$

where K_c is equilibrium constant

In a dilute solution the concentration of the water is not going to change significantly during the dissociation process, and so for these reactions concentration of water can be taken as constant, hence the expression for acid dissociation can be written where K_a is the acid dissociation constant.

An indicator is in fact a weak acid, HIn. Like other weak acids, it dissociates, so it forms H+ and In- ions. The HIn will be one colour and the In- will be a different colour.

HIn H+   +   In-

Colour 1   Colour 2

- ✓ In an acid: high concentration of H+ → equilibrium to move to the left→ colour 1.
- ✓ In alkali: H+ will be removed →equilibrium will shift to the right →colour 2.

When the amount of HIn and are exactly balanced the colour will be in between the two colours. One drop of acid or alkali, and the equivalence point should be able to change the colour of the indicator.

The **end poin**t found as a weak acid (K_a, or K_{In}) → end point takes place at a pH equal to its pK_{In} value

Indicator	pH range	Color		pK_{In}
		Acid	Alkaline	
Thymol Blue (acid)	1.2 – 2.8	Red	Yellow	1.7
Methyl Orange	2.9 – 4.6	Pink	Orange	3.7
Methyl Red	4.2 – 6.3	Red	Yellow	5.0
Bromothymol Blue	6.0 – 7.6	Yellow	Blue	7.1
Phenolphthalein	8.3 – 10.0	Colorless	Pink	9.6
Thymolphthalein	9.3 – 10.5	Colorless	Blue	9.3

Titration Curves and selecting and indicator

End point = mid way between 2 colours of indicator (a property of the indicator).
Equivalence point = when the stoichiometric amounts of acid and alkali have been added.
End point and equivalence point must coincide for an effective titration.

If the indicator is chosen correctly the end point and the equivalence point are the same.
A pH meter, or conductimetric method, can also be used to determine the equivalence point in an acid/base titration.

The titration curve serves to profile the unknown solution. The shape of the curve lies much in the properties of acid and base used.

You need to choose an **indicator** which **changes colour as close as possible to that equivalence point**. That varies from titration to titration.