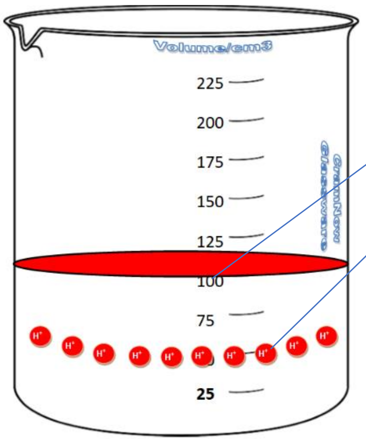
The Neutralisation of Strong Acid by Strong Base and the Titration Curve

For this, we will consider how the pH changes as we progressively neutralise hydrochloric acid by adding sodium hydroxide.

Let's begin by having **100cm^{-3} of hydrochloric in a beaker**. The concentration is 0.100 mol dm^{-3}

This is represented by the following diagram

In these diagrams H^+ represents 1 x 10^{-3} mol **dissociated** H^+ ions that **contribute to the pH.**



volume of solution = 0.100 dm^{-3}

moles of H^+, n_{H^+} = 10.0 x 10^{-3} mol

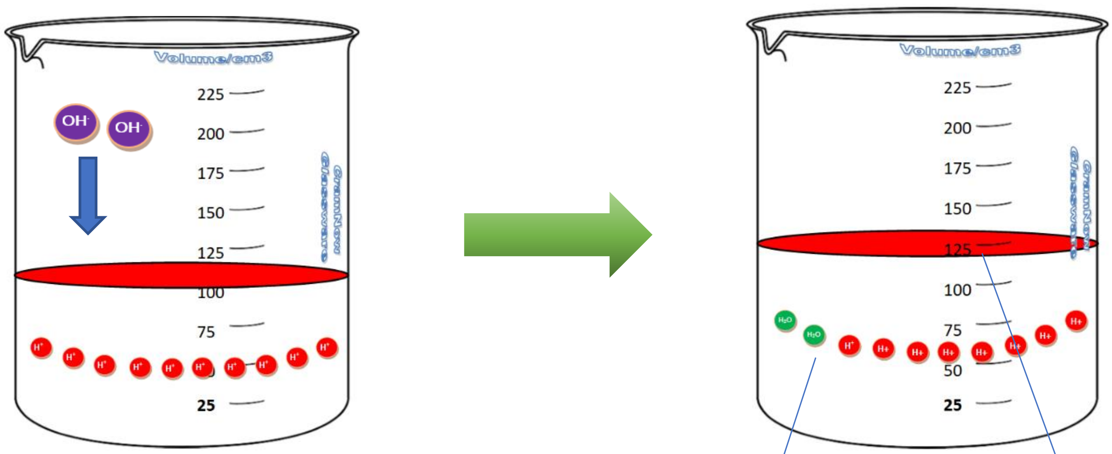
$$[H^+] = \frac{10.0 \times 10^{-3}\ \text{mol}}{0.100\ \text{dm}^{-3}} = 1.00 \times 10^{-1}\ \text{mol dm}^{-3}$$

pH = -log [H$^+$] = -log 1.0 x 10^{-1} = 1.00 = pH

Let's now add 20cm^3 of 0.100 mol dm^{-3} NaOH solution. This will increase the volume by 20cm^3.

OH^- Represents 1 x 10^{-3} mol OH^- ions that are added to the beaker.

As the OH^- ions meet the H^+ ions, there is a reaction that produces H_2O

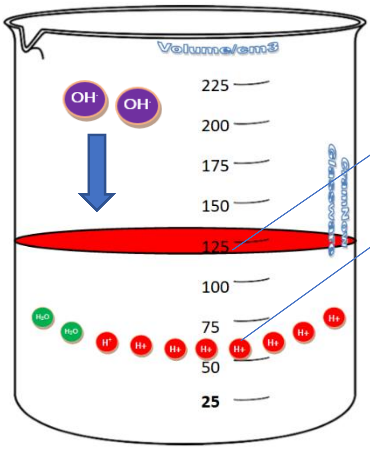


Two obvious things have happened.

The number of H^+ **ions has been reduced** and the **volume has increased.**

It seems logical that **both changes** must make the **solution less acidic** and **increase the pH.**

We need to be able to calculate the new pH. This is done below.



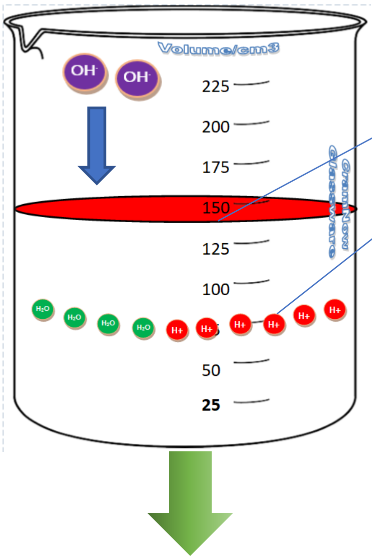
new volume = 0.120 dm^{-3} (100cm^3 + 20cm^3!)

new n_{H^+} = n_{H^+} before addition – n_{OH^-} added
= 10 x 10^{-3} mol – 2.0 x 10^{-3} mol
= 8 x 10^{-3} mol of H^+

$[H^+]$ = $\frac{8 \times 10^{-3}\ \text{mol}}{0.120\ \text{dm}^{-3}}$ = 6.67 x 10^{-2} mol dm^{-3}

pH = -log $[H^+]$ = -log 6.67 x 10^{-2} = 1.18 = pH

Let's repeat this by adding a **further 20cm^3 of 0.100 mol dm^{-3} NaOH solution**. You do the calculations.



new volume =

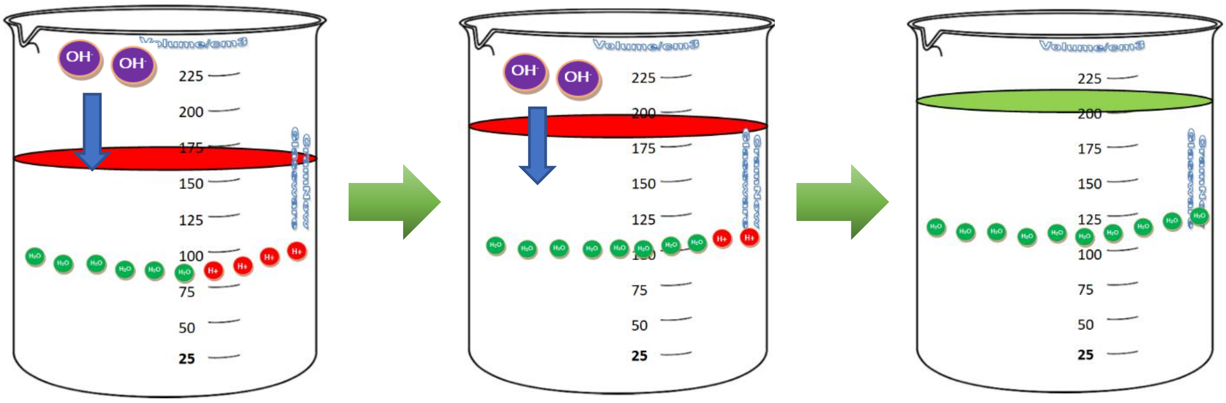
new n_{H^+} = n_{H^+} before addition – n_{OH^-} added

= .... - ....

= ....mol of H^+

new $[H^+]$ = ____________ =...................mol dm^{-3}

pH = -log $[H^+]$ = -log .... = ... = pH



new volume	= ....
new n_{H^+}	= ....
$[H^+]$	=...............
pH	=...............

new volume	= ....
new n_{H^+}	= ....
$[H^+]$	=...............
pH	=...............

new volume	= ....
new n_{H^+}	= care! Not 0!
$[H^+]$	=think!........
pH	=think!........

Fill in the blank boxes in the table below.

ADDITION of OH^- UP TO EQUIVALENCE POINT

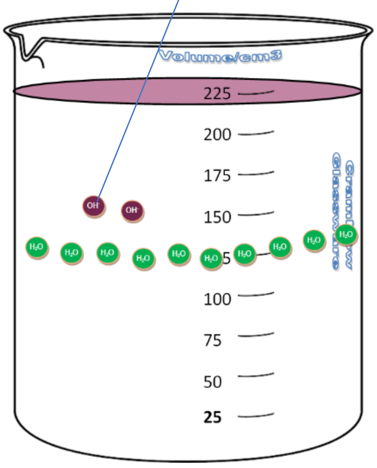
total volume of 0.100 M NaOH added/cm^3	0.0	20.0	40.0	60.0	80.0	90.0	95.0	99.0	99.5	99.9	100
pH	1.00	1.18	1.37	1.60	1.95						equiv. point

These are the answers that you should have calculated on the previous page

The **equivalence point** where all of the hydrochloric acid has been neutralised. There is only a neutral salt solution at this point.

We need to consider what would happen to the pH if we continue to add more OH^- ions beyond the equivalence point.
If we pour in another 20cm^3 of 0.100 mol dm^{-3} NaOH solution, the situation will be the one in the diagram below.

Excess OH^- ions that have no H^+ (from the original acid) to react with.
This has **increased the OH^- ion concentration and increased the volume.**



new volume = 0.220 dm^{-3} (200cm^3 + 20cm^3!)

n_{OH^-} added = (20/1000) x 0.100 mol dm^{-3}
= 2.0×10^{-3} mol of OH^-

$$[OH^-] = \frac{2.0 \times 10^{-3} \text{ mol}}{0.220 \text{ dm}^{-3}} = 9.09 \times 10^{-3} \text{ mol dm}^{-3}$$

$$[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{9.09 \times 10^{-3}} = 1.10 \times 10^{-11}$$

pH = -log [H^+] = -log 1.10×10^{-11} = 10.9 = pH

If we continued adding 20cm^3 samples of 0.100 mol dm^{-3} NaOH to this solution, the concentration of OH^- **would continue to grow and so would the volume (we'd need a larger beaker!).**

After a further 20cm^3 samples of 0.100 mol dm^{-3} NaOH solution has been added, the pH will be:

new volume = **0.240 dm^{-3}** (220cm^3 + **20cm^3!**)

n_{OH^-} added = (20/1000) x 0.100 mol dm^{-3}

= **2.0 x 10^{-3}** mol of OH^-

new n_{OH^-} = **4.0 x 10^{-3}** mol of OH^-

$$[OH^-] = \frac{4.0 \times 10^{-3}\ \text{mol}}{0.240\ \text{dm}^{-3}} = \mathbf{1.67 \times 10^{-2}}\ \text{mol dm}^{-3}$$

$$[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{1.67 \times 10^{-2}} = \mathbf{6.00 \times 10^{-13}}\ \text{mol dm}^{-3}$$

pH = -log [H^+] = -log **6.00 x 10^{-13}** = 12.2 = pH

Below is a continuation of the table from earlier. Fill in the missing blanks.

ADDITION of OH^- BEYOND EQUIVALENCE POINT

total volume of 0.100 M NaOH added/cm^3	100	100.1	100.5	101.0	105.0	110.0	120.0	140.0	160.0	180.0	200
pH	7.00 equiv. point						10.9	12.2			

You now should have a full set of pHs for a large range of volumes of 0.100 mol dm^{-3} NaOH solution from 20.0cm^3 to 200cm^3.

This data should now be plotted on graph paper with volume of NaOH solution on the x axis and pH on the y axis.

This will produce a pH titration curve for a STRONG ACID (hydrochloric acid) 'versus' STRONG BASE (sodium hydroxide)

Graph paper is on the following page

STRONG ACID (hydrochloric acid) 'versus' STRONG BASE (sodium hydroxide)