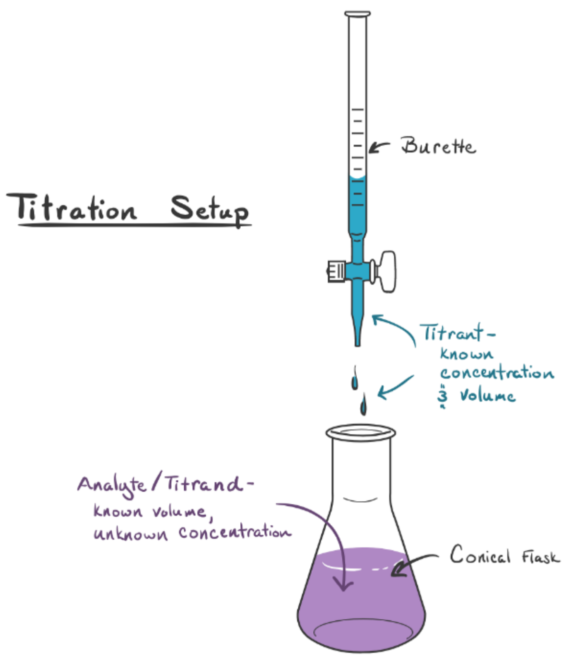
Titration

Method used to determine the concentration of a solution.

There are two solutions used in a titration.

Solution	Location	Is its Concentration Known Before Titration Begins?	Is its Volume Known Before Titration Begins?
Analyte	Flask	No! This is what we are ultimately solving for.	Yes! We must know the exact volume of analyte placed into the flask.
Titrant	Buret	Yes! If we didn't know this, then the titration would be pointless.	No! This quantity will be measured by doing the experiment. Add titrant to the flask until the rxn reaches the equivalence point. Record volume of titrant used.



Equivalence point: stoichiometrically **equal moles of analyte and titrant (i.e. moles of acid = moles of base)**

End point: when indicator **changes color**, signifies the end of titration. We want to use an indicator that will change color close to the equivalence point.

Indicators: use an indicator that will change color in the pH range of the equivalence point.

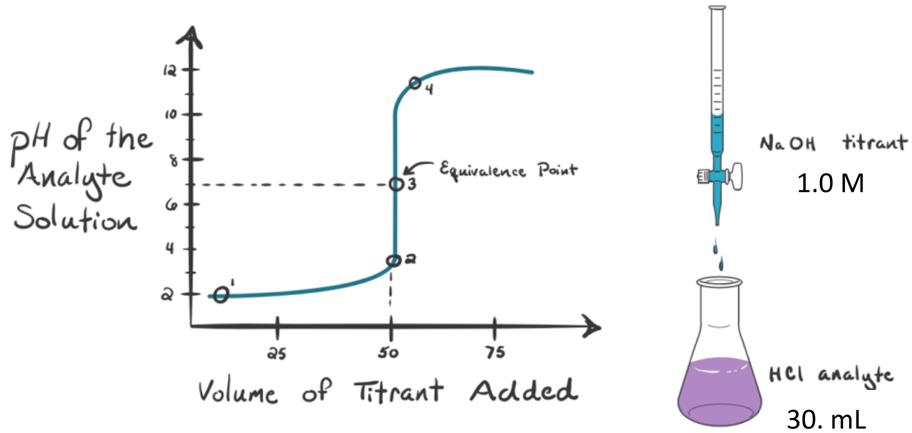
- Choose an indictor whose pKa = pH of the solution at equivalence point
- Example: if pH = 4 at equivalence point, choose an indicator with a pKa = 4 (i.e. $K_a = 1 \times 10^{-4}$)

3 Types of Acid Base Reactions:

Rxn Type	**Strong Acid + Strong Base**	**Weak Acid + Strong Base**	**Strong Acid + Weak Base**
Net Ionic Equation	$H^+ + OH^- \rightarrow H_2O$	$HA + OH^- \rightarrow H_2O + A^-$	$H^+ + B \rightarrow BH^+$
Products	Water + Neutral Salt	Water + Basic Salt	Water + Acidic Salt
pH at Equivalence Point	pH = 7	pH > 7	pH < 7
Example	$HCl + NaOH \rightarrow H_2O + NaCl$	$HF + NaOH \rightarrow H_2O + NaF$	$HCl + NH_3 \rightarrow NH_4Cl$
	Cross out any ions part of a strong acid or strong base (because they are spectators). You got your net ionic equation! 😊		

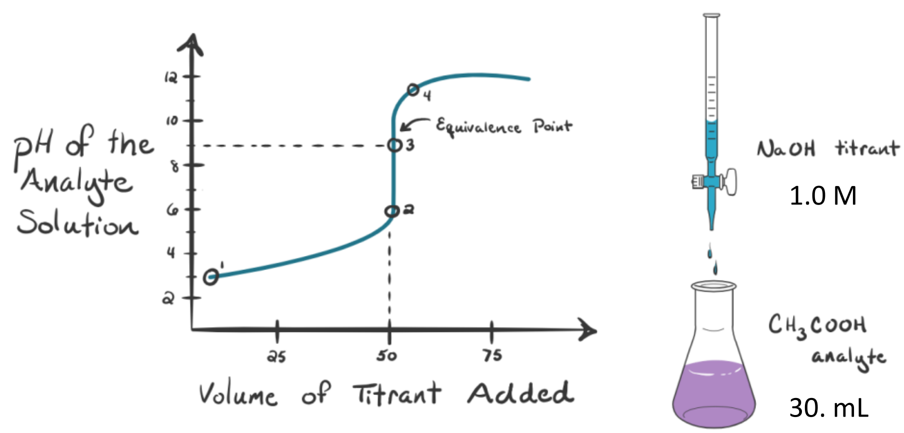
Titration curve: plot of pH (y-axis) vs. volume of titrant added (x-axis)

Example 1



Rxn type	
Net ionic equation	
pH at equivalence point	
Volume of titrant needed to reach equivalence point	
Calculate concentration of analyte	

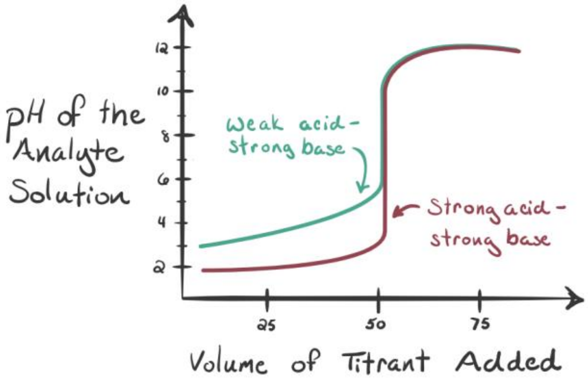
Example 2

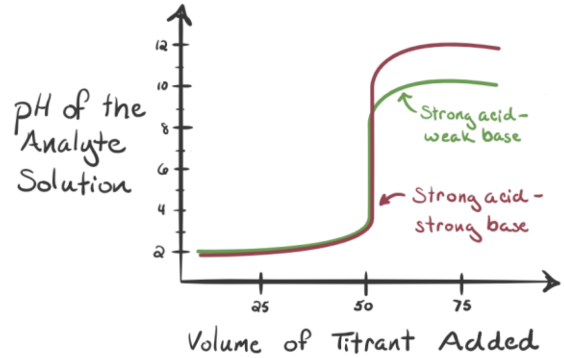


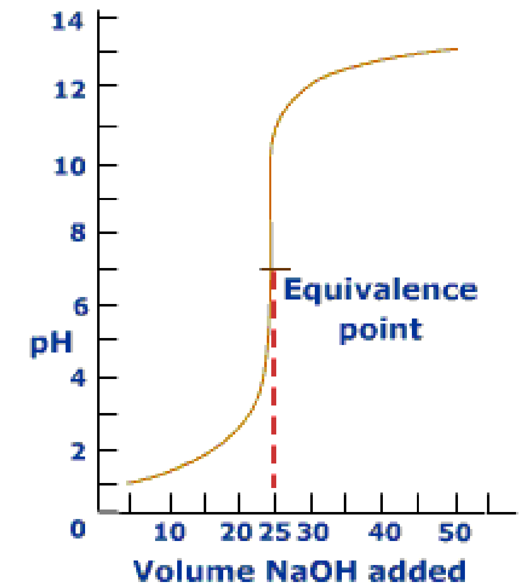
Rxn type	
Net ionic equation	
pH at equivalence point	
Volume of titrant needed to reach equivalence point	
Calculate concentration of analyte	

How do you use a titration curve to determine the equivalence point?

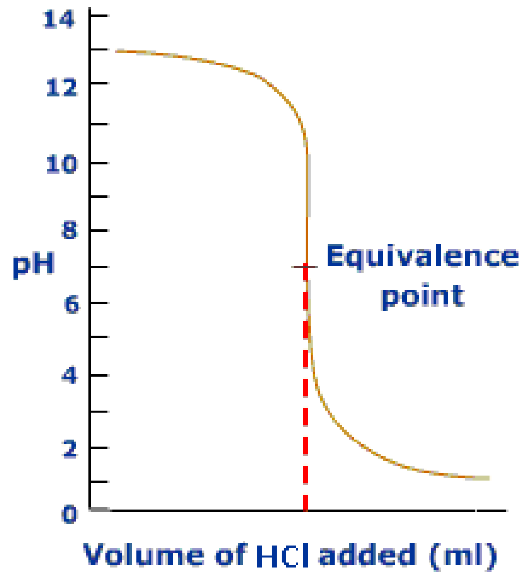
Why is the pH at the equivalence point different for the two examples above?



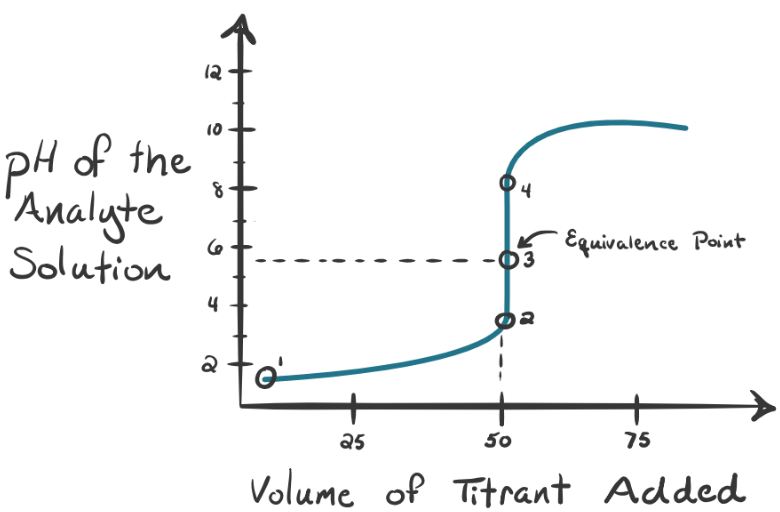




Titration curve of strong acid (HCl) with a strong base (NaOH)



Titration curve of strong base (NaOH) with strong acid (HCl)



What type of reaction is represented by the titration curve to the left?

5 Main Regions of an Acid Base Titration Curve

Strong Acid + Strong Base Titration

		What's in the flask?	How to calculate pH
1.	Before titration begins	Analyte only	Initial [analyte] = $[H^+]$ or $[OH^-]$ pH = – $\log[H^+]$ pOH = – log $[OH^-]$ pH + pOH = 14
2.	Before equivalence point	Fewer moles of analyte	$[H^+]$ or $[OH^-]$ = $\left(\frac{mols\ analyte\ remaing}{total\ volume}\right)$
3.	Half-way to equivalence point	Half the moles of original analyte	$[H^+]$ or $[OH^-]$ = $\left(\frac{mols\ analyte\ remaing}{total\ volume}\right)$
4.	Equivalence point	Water and salt Zero moles of analyte	Neutral salt pH = 7
5.	Beyond equivalence point	Excess titrant	$[H^+]$ or $[OH^-]$ = $\left(\frac{mols\ titrant\ in\ excess}{total\ volume}\right)$

Weak Acid/Base + Strong Base/Acid Titration

		What's in the flask?	How to calculate pH
1.	Before titration begins	Analyte only	Weak acid = Ka ICE chart Weak base = Kb ICE chart
2.	Before equivalence point	Buffer	$[H^+]$ = Ka $\left(\frac{mols\ weak\ acid}{mols\ conjugate\ base}\right)$
3.	Half-way to equivalence point	Buffer with 1:1 mole ratio of weak acid to conjugate base	pH = pKa of weak acid
4.	Equivalence point	Water and salt (i.e. only conjugate acid/base of the analyte)	Acidic salt = Ka ICE chart, pH < 7 Basic salt = Kb ICE chart, pH > 7
5.	Beyond equivalence point	Excess titrant	$[H^+]$ or $[OH^-]$ = $\left(\frac{mols\ titrant\ in\ excess}{total\ volume}\right)$