Strong acid/strong base titration

10.0 mL of 0.50 *M* HCl is titrated with 0.10 *M* NaOH. Calculate the **pH** at the following points in the titration.

a) 0.00 mL of NaOH added

At the start of the titration only strong acid, HCl, is present. The concentration of the HCl solution is 0.50 *M*

	$HCl + H_2O \rightarrow$	H_3O^+ +	Cl^-
Initial (*M*)	0.50	0	0
Change (*M*)	–0.50	+0.50	+0.50
After Rxn (*M*)	0	0.50	0.50

pH = $-\log[H_3O^+] = -\log(0.50) =$ **0.30**

b) 10.0 mL of NaOH added

Acid reacting with base typically produces water and a salt. Let's initially work with moles of acid and base since volume is changing. Remember, $M \times L = mol$. After the reaction is complete, we will convert back to molarity to calculate the pH.

	HCl +	$NaOH \rightarrow H_2O + NaCl$
Initial (mol)	0.0050	0.0010
Change (mol)	–0.0010	–0.0010
After Rxn (mol)	0.0040	0

After the reaction, 0.0040 mol of HCl remain. Let's convert to molarity of HCl. The total volume of the solution is 10 mL HCl + 10 mL NaOH = 20 mL solution.

$$M = \frac{0.0040 \text{ mol}}{0.020 \text{ L}} = 0.20\ M$$

This is the concentration of the strong acid, HCl.

	$HCl + H_2O \rightarrow$	H_3O^+ +	Cl^-
Initial (*M*)	0.20	0	0
Change (*M*)	–0.20	+0.20	+0.20
After Rxn (*M*)	0	0.20	0.20

pH = $-\log[H_3O^+] = -\log(0.20) =$ **0.70**

c) 30.0 mL of NaOH added

Again, let's work with moles of acid and base since volume is changing. We will convert back to molarity after the reaction is complete.

	HCl +	$NaOH \rightarrow H_2O + NaCl$
Initial (mol)	0.0050	0.0030
Change (mol)	–0.0030	–0.0030
After Rxn (mol)	0.0020	0

After the reaction, 0.0020 mol of HCl remain. Let's convert to molarity of HCl. The total volume of the solution is 10 mL HCl + 30 mL NaOH = 40 mL solution.

$$M = \frac{0.0020 \text{ mol}}{0.040 \text{ L}} = 0.050\ M$$

This is the concentration of the strong acid, HCl.

	$HCl + H_2O \rightarrow H_3O^+$		+ Cl^-
Initial (M)	0.050	0	0
Change (M)	–0.050	+0.050	+0.050
After Rxn (M)	0	0.050	0.050

pH = $-\log[H_3O^+] = -\log(0.050) =$ **1.30**

d) 50.0 mL of NaOH added

Again, let's work with moles of acid and base since volume is changing. We will convert back to molarity after the reaction is complete.

	HCl +	$NaOH \rightarrow H_2O + NaCl$
Initial (mol)	0.0050	0.0050
Change (mol)	–0.0050	–0.0050
After Rxn (mol)	0	0

This is the equivalence point of the titration. Just enough base has been added to completely reaction with the acid. For a strong acid/strong base titration the pH at the equivalence point is **7**. Why is this? The only species present in the solution at the equivalence point is the salt, NaCl. Will this salt change the pH of water?

e) 65.0 mL of NaOH added

	HCl +	$NaOH \rightarrow H_2O + NaCl$
Initial (mol)	0.0050	0.0065
Change (mol)	–0.0050	–0.0050
After Rxn (mol)	0	0.0015

After the reaction, 0.0015 mol of NaOH remain. We are now past the equivalence point. Excess base is accumulating in the solution. Let's convert to molarity of NaOH. The total volume of the solution is 10 mL HCl + 65 mL NaOH = 75 mL solution.

$$M = \frac{0.0015 \text{ mol}}{0.075 \text{ L}} = 0.020\ M$$

This is the concentration of the strong base, NaOH.

	$NaOH + H_2O \rightarrow Na^+$		+ OH^-
Initial (M)	0.020	0	0
Change (M)	–0.020	+0.020	+0.020
After Rxn (M)	0	0.020	0.020

pOH = $-\log[OH^-] = -\log(0.020) = 1.70$
pH = 14 – pOH = **12.30**