

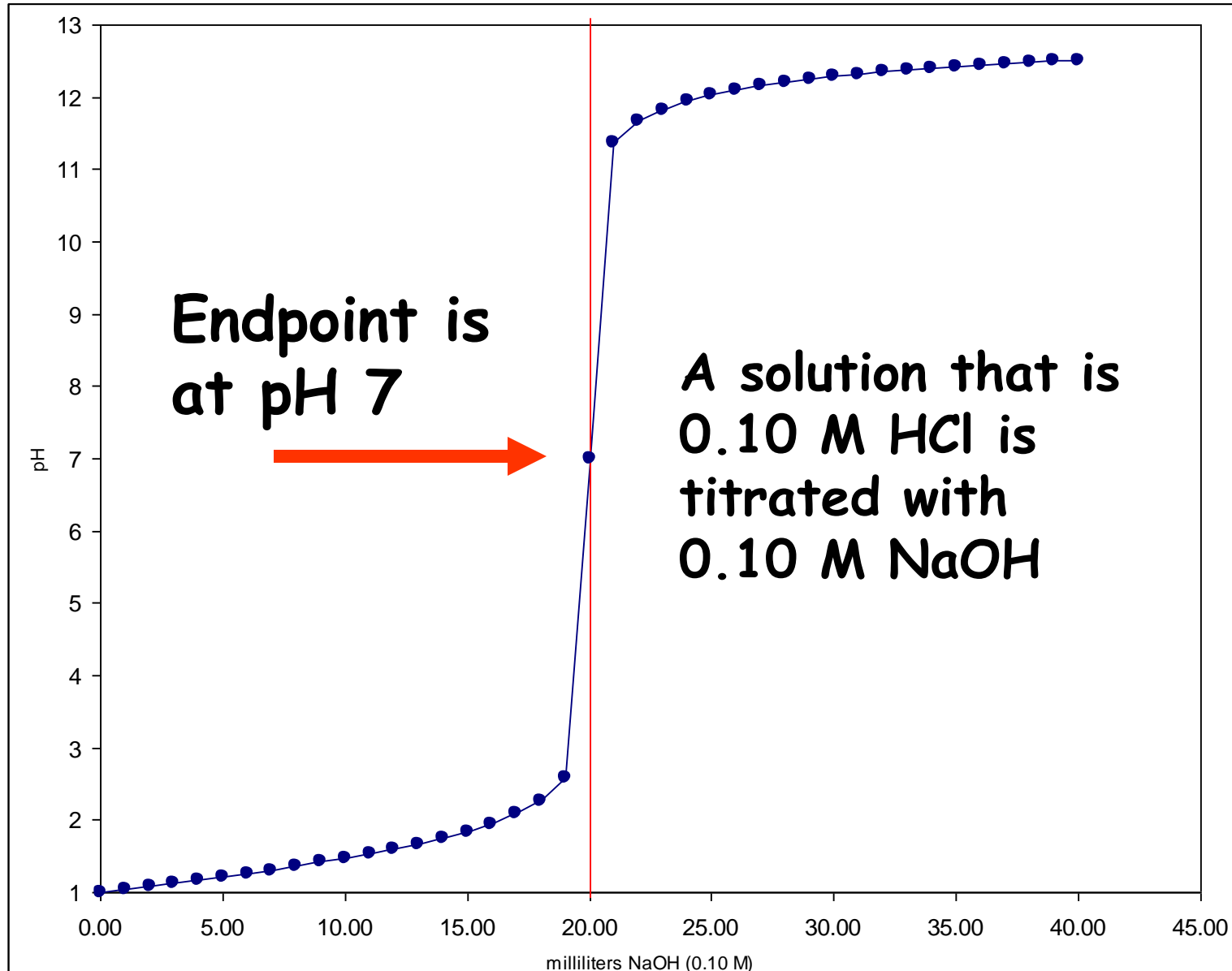
# **N40 – Acid Base**

**Target: I can perform titration calculations.**

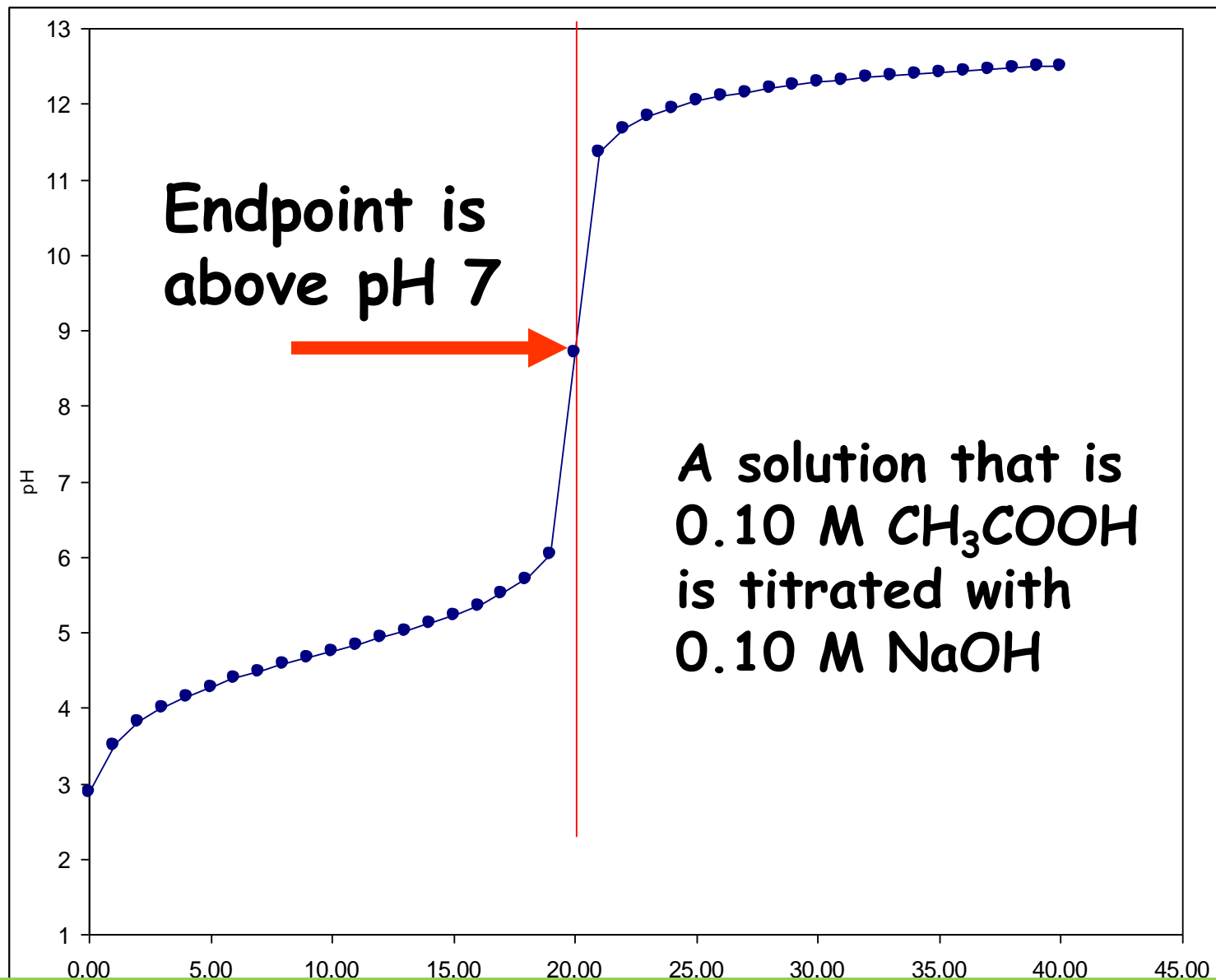
# **N40 – Acid Base**

## **Titration**

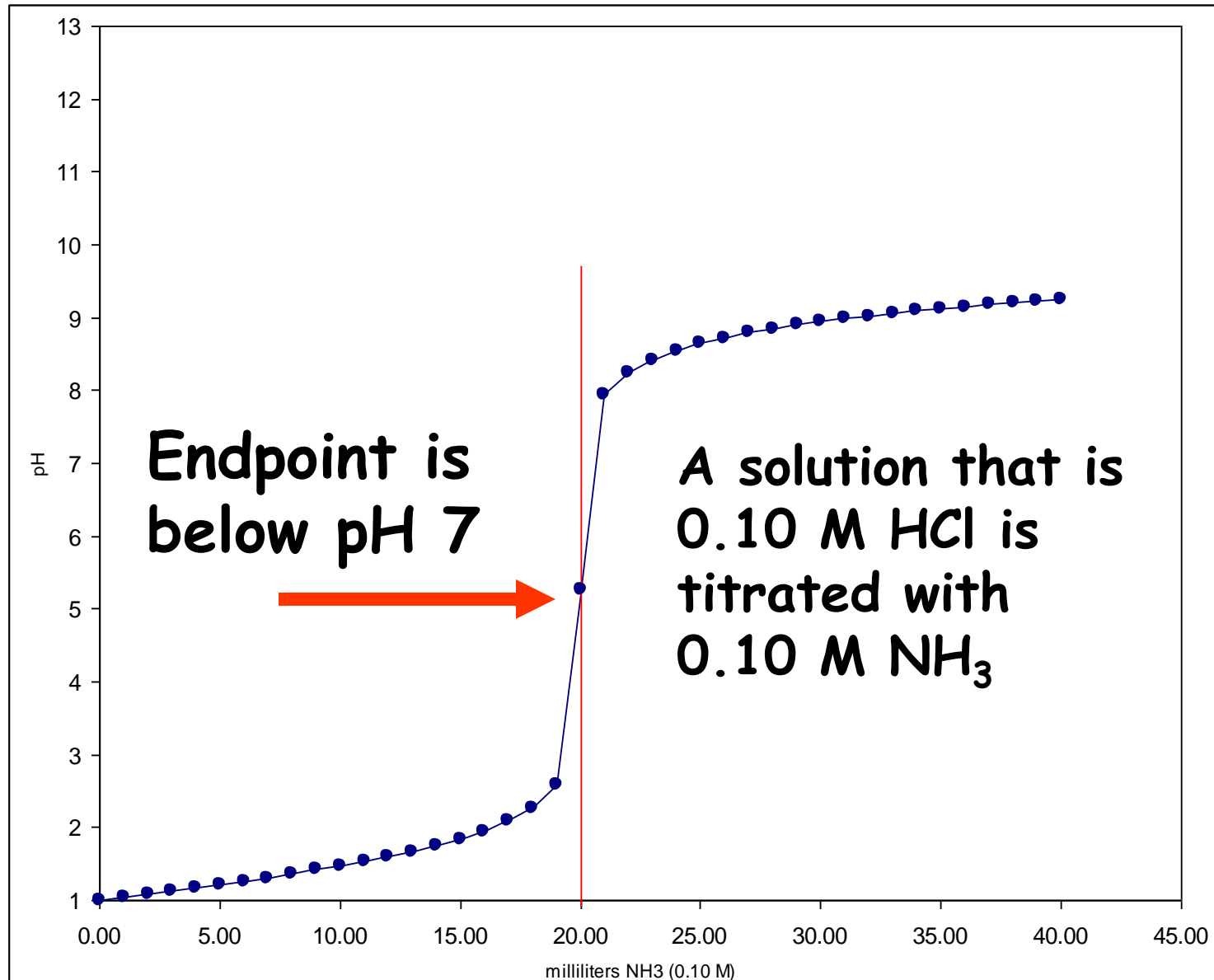
# Strong Acid/Strong Base Titration



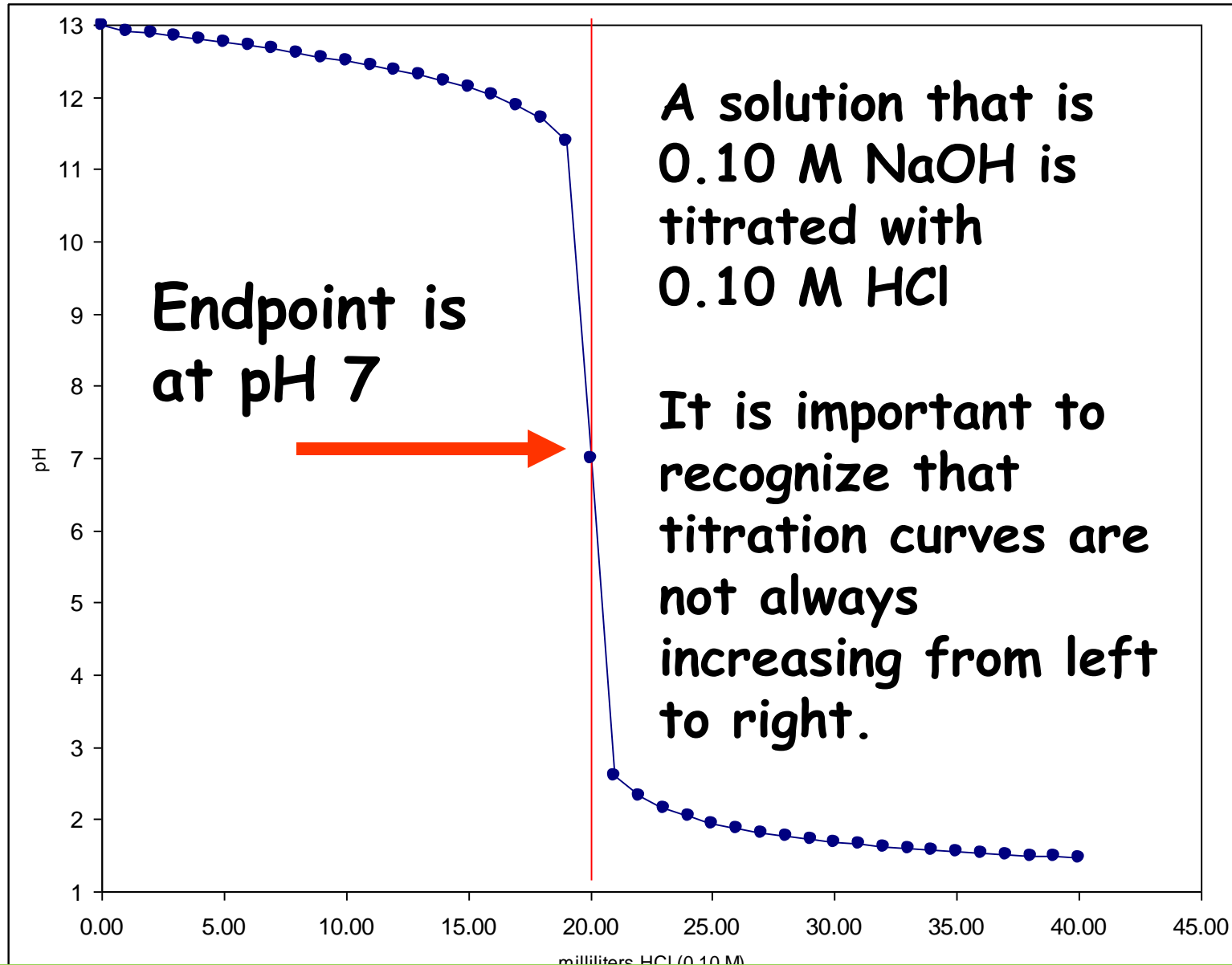
# Weak Acid/Strong Base Titration



# Strong Acid/Weak Base Titration



# Strong Acid/Strong Base Titration



# Titration Calculations...

## 1. Starting pH

- ICE table then pH

## 2. Early on during titration

- Stoich then He-Ha

## 3. Equivalence Point

- mol acid = mol base
- No more buffer! Reverse rxn
- Calc new K value - ICE then pH

## 4. $\frac{1}{2}$ Way Point

- $\frac{1}{2}$  moles @ eq.pt
- pH = pKa

## 5. Towards end of titration

- Extra titrant left over
- Stoich then simple pH

**BRACE YOURSELF**



# Calculations to Plot a Titration Curve

## 1. Starting pH

- ICE table then pH

## 2. Early on during titration

- Stoich then He-Ha

## 3. Equivalence Point

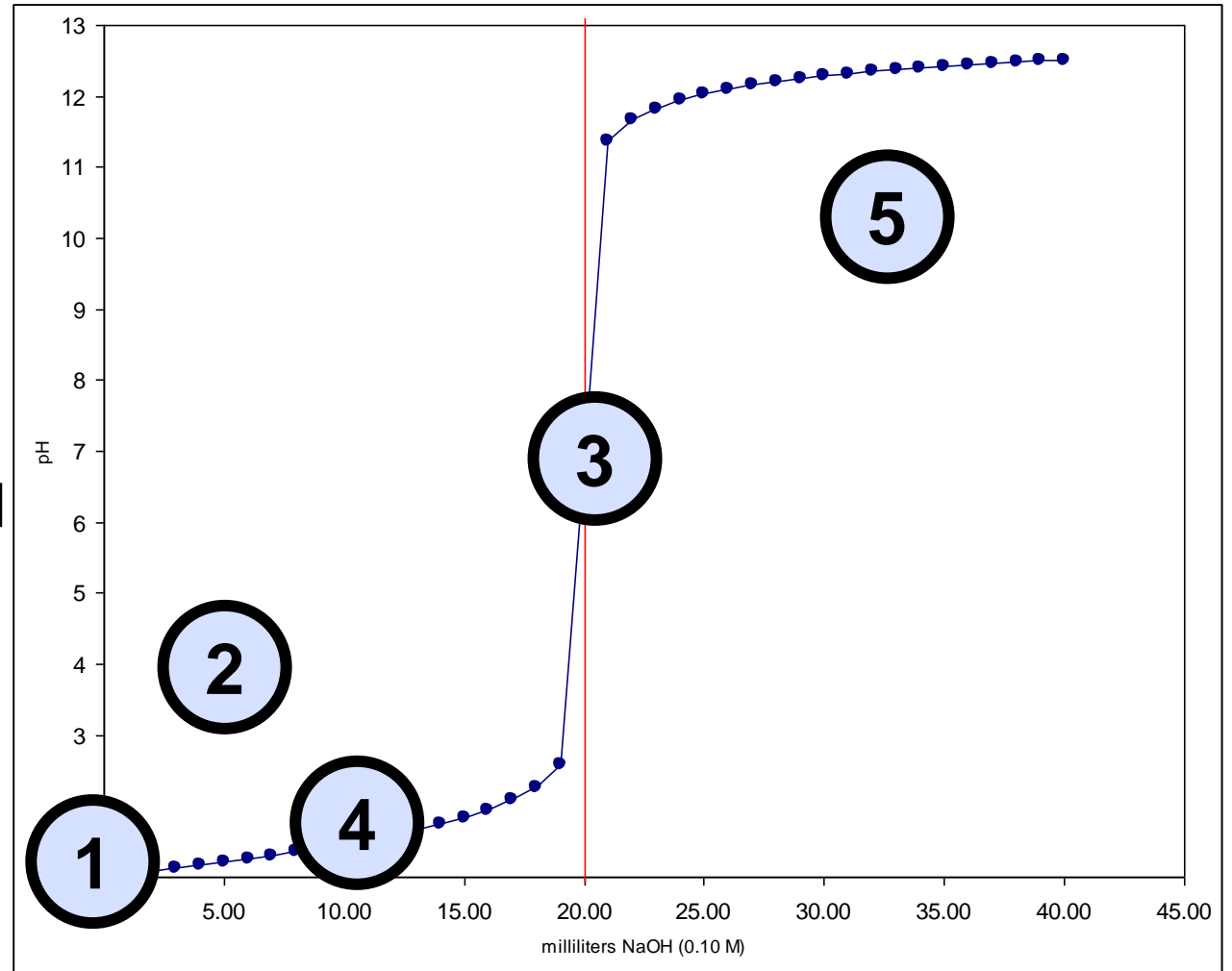
- mol acid = mol base
- No more buffer! Reverse rxn
- Calc new K value - ICE then pH

## 4. $\frac{1}{2}$ Way Point

- $\frac{1}{2}$  moles @ eq.pt
- pH = pKa

## 5. Towards end of titration

- Extra titrant left over
- Stoich then simple pH



# Titration Calculations

1

Lets look at the titration of acetic acid w/ NaOH

## BEFORE TITRATION

- Starting point:
  - 25 ml of 0.15M Acetic Acid ( $K_a = 1.8E^{-5}$ )
  - Calculate pH before any titrant is added
    - ICE TABLE! Then pH calculation

$C_2H_3OH$	$\leftrightarrow$	$H^+$	$C_2H_3O^-$
0.15 M		0 M	0 M
- x		+ x	+ x
$0.15 - x$		x	x
0.15		x	x

$$K = \frac{[H^+][C_2H_3O^-]}{[C_2H_3OH]}$$

$$1.8 \times 10^{-5} = \frac{(x)(x)}{0.15}$$

$$x = 1.64 \times 10^{-3}$$

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

$$pH = -\log(1.64 \times 10^{-3})$$

$$pH @ \text{starting point} \\ = 2.78$$

# Titration Calculations

2

## DURING THE TITRATION

- Add 10ml of 0.10 M NaOH
  - Determine stoichiometry
  - Notice! You have a buffer now.
    - You have a conjugate base!
    - Use He-Ha eq.

$$\frac{25 \text{ mL}}{1000 \text{ mL}} \times \frac{0.15 \text{ mol}}{1 \text{ L}} = 3.75 \times 10^{-3} \text{ mol acid}$$

$$\frac{10 \text{ mL}}{1000 \text{ mL}} \times \frac{0.10 \text{ mol}}{1 \text{ L}} = 1.0 \times 10^{-3} \text{ mol base}$$

People like to use these “mole tables” – they are NOT ICE TABLES! They have moles not concentrations. **BE CAREFUL!**

$\text{C}_2\text{H}_3\text{OH}$	$\text{OH}^- \leftrightarrow \text{H}_2\text{O}$	$\text{C}_2\text{H}_3\text{O}^-$
3.75 mmol	1 mmol	0
2.75 mmol	0 mmol	1 mmol

**Have to convert to M before using He-Ha!**

$$[\text{C}_2\text{H}_3\text{OH}] = \frac{2.75 \times 10^{-3} \text{ mol}}{(0.025 \text{ L} + 0.010 \text{ L})} = 0.0786 \text{ M}$$

$$[\text{C}_2\text{H}_3\text{O}^-] = \frac{1.00 \times 10^{-3} \text{ mol}}{(0.025 \text{ L} + 0.010 \text{ L})} = 0.0286 \text{ M}$$

**NOW you can use He-Ha**

# Titration Calculations

2

## DURING THE TITRATION

- Add 10ml of 0.10 M NaOH
  - Determine stoichiometry
  - Notice! You have a buffer now.
    - You have a conjugate base!
    - Use He-Ha eq.

$$[C_2H_3OH] = \frac{2.75 \times 10^{-3} \text{ mol}}{(0.025 \text{ L} + 0.010 \text{ L})} = 0.0786 \text{ M}$$

$$[C_2H_3O^-] = \frac{1.00 \times 10^{-3} \text{ mol}}{(0.025 \text{ L} + 0.010 \text{ L})} = 0.0286 \text{ M}$$

$$pH = pKa + \log \left( \frac{[A^-]}{[HA]} \right)$$

$$pH = -\log(1.8 \times 10^{-5}) + \log \left( \frac{0.0286 \text{ M}}{0.0786 \text{ M}} \right)$$

$$pH = 4.31$$

# Titration Calculations

2

## DURING THE TITRATION AGAIN

- Add 25ml of 0.10 M NaOH
  - Determine stoichiometry
  - Notice! You have a buffer now.
    - You have a conjugate base!
    - Use He-Ha eq.

$$\frac{25 \text{ mL}}{1000 \text{ mL}} \times \frac{0.15 \text{ mol}}{1 \text{ L}} = 3.75 \times 10^{-3} \text{ mol acid}$$

$$\frac{25 \text{ mL}}{1000 \text{ mL}} \times \frac{0.10 \text{ mol}}{1 \text{ L}} = 2.5 \times 10^{-3} \text{ mol base}$$

People like to use these “mole tables” – they are NOT ICE TABLES! They have moles not concentrations. **BE CAREFUL!**

$\text{C}_2\text{H}_3\text{OH}$	$\text{OH}^- \leftrightarrow \text{H}_2\text{O}$	$\text{C}_2\text{H}_3\text{O}^-$
3.75 mmol	2.5 mmol	0
1.25 mmol	0 mmol	2.5 mmol

**Have to convert to M before using He-Ha!**

$$[\text{C}_2\text{H}_3\text{OH}] = \frac{1.25 \times 10^{-3} \text{ mol}}{(0.025 \text{ L} + 0.025 \text{ L})} = 0.025 \text{ M}$$

$$[\text{C}_2\text{H}_3\text{O}^-] = \frac{1.00 \times 10^{-3} \text{ mol}}{(0.025 \text{ L} + 0.025 \text{ L})} = 0.050 \text{ M}$$

**NOW you can use He-Ha**

# Titration Calculations

2

## DURING THE TITRATION AGAIN

- Add 25ml of 0.10 M NaOH
  - Determine stoichiometry
  - Notice! You have a buffer now.
    - You have a conjugate base!
    - Use He-Ha eq.

$$[C_2H_3OH] = \frac{1.25 \times 10^{-3} \text{ mol}}{(0.025 \text{ L} + 0.025 \text{ L})} = 0.025 \text{ M}$$

$$[C_2H_3O^-] = \frac{1.00 \times 10^{-3} \text{ mol}}{(0.025 \text{ L} + 0.025 \text{ L})} = 0.050 \text{ M}$$

$$pH = pKa + \log \left( \frac{[A^-]}{[HA]} \right)$$

$$pH = -\log(1.8 \times 10^{-5}) + \log \left( \frac{0.050 \text{ M}}{0.025 \text{ M}} \right)$$

$$pH = 5.05$$

# Titration Calculations

3

## AT EQUIVALENCE POINT

- Add 37.5 ml of 0.10 M NaOH
  - Determine stoichiometry
  - Notice! You have NO BUFFER LEFT!
    - You have NO weak acid left!!

People like to use these “mole tables” – they are NOT ICE TABLES! They have moles not concentrations. **BE CAREFUL!**

$\text{C}_2\text{H}_3\text{OH}$	$\text{OH}^-$	$\longleftrightarrow \text{H}_2\text{O}$	$\text{C}_2\text{H}_3\text{O}^-$
3.75mmol	3.75mmol	0	0
0 mmol	0 mmol	3.75mmol	3.75mmol

**Have to Reverse the Rxn, new ICE table!**

$$\frac{25 \text{ mL}}{1000 \text{ mL}} \times \frac{0.15 \text{ mol}}{1 \text{ L}} = 3.75 \times 10^{-3} \text{ mol acid}$$

$$\frac{37.5 \text{ mL}}{1000 \text{ mL}} \times \frac{0.10 \text{ mol}}{1 \text{ L}} = 3.75 \times 10^{-3} \text{ mol base}$$

# Titration Calculations

3

## AT EQUIVALENCE POINT

- Add 37.5 ml of 0.10 M NaOH
  - Determine stoichiometry
  - Notice! You have NO BUFFER LEFT!
    - You have NO weak acid left!!

*Remember to use M in ICE Table not moles!*

$$[C_2H_3O^-] = \frac{3.75 \times 10^{-3} \text{ mol}}{(0.025 \text{ L} + 0.0375 \text{ L})} = 0.060 \text{ M}$$

People like to use these “mole tables” – they are NOT ICE TABLES! They have moles not concentrations. **BE CAREFUL!**

$C_2H_3OH$	$OH^-$	$\leftrightarrow H_2O$	$C_2H_3O^-$
3.75mmol	3.75mmol	0	0
0 mmol	0 mmol	3.75mmol	3.75mmol

*Have to Reverse the Rxn, new ICE table!*

$C_2H_3O^-$	$H_2O$	$\leftrightarrow C_2H_3OH$	$OH^-$
0.06 M	-	0	0
- x	-	+ x	+ x
0.06	-	x	x

# Titration Calculations

3

## AT EQUIVALENCE POINT

- Add 37.5 ml of 0.10 M NaOH
  - Determine stoichiometry
  - Notice! You have NO BUFFER LEFT!
    - You have NO weak acid left!!

*Remember to use  $K_b$  this time!*

$$K_w = K_a \times K_b \quad K_b = \frac{K_w}{K_a}$$

$$K_b = \frac{(1 \times 10^{-14})}{(1.8 \times 10^{-5})} = 5.56 \times 10^{-10}$$

People like to use these “mole tables” – they are NOT ICE TABLES! They have moles not concentrations. **BE CAREFUL!**

$C_2H_3OH$	$OH^-$	$\leftrightarrow H_2O$	$C_2H_3O^-$
3.75mmol	3.75mmol	0	0
0 mmol	0 mmol	3.75mmol	3.75mmol

*Have to Reverse the Rxn, new ICE table!*

$C_2H_3O^-$	$H_2O$	$\leftrightarrow C_2H_3OH$	$OH^-$
0.06 M	-	0	0
- x	-	+ x	+ x
0.06	-	x	x

# Titration Calculations

3

## AT EQUIVALENCE POINT

- Add 37.5 ml of 0.10 M NaOH
  - Determine stoichiometry
  - Notice! You have NO BUFFER LEFT!
    - You have NO weak acid left!!

*Remember to use  $K_b$  this time!*

$$K_w = K_a \times K_b \quad K_b = \frac{K_w}{K_a}$$

$$K_b = \frac{(1 \times 10^{-14})}{(1.8 \times 10^{-5})} = 5.56 \times 10^{-10}$$

$C_2H_3O^-$	$H_2O \leftrightarrow C_2H_3OH$	$OH^-$
0.06 M	-	0
- x	-	+ x
0.06	-	x

$$5.56 \times 10^{-10} = \frac{(x)(x)}{0.06}$$

$$x = 5.77 \times 10^{-6} = [OH^-]$$

*Now you can do pH calculation!*

# Titration Calculations

3

## AT EQUIVALENCE POINT

- Add 37.5 ml of 0.10 M NaOH
  - Determine stoichiometry
  - Notice! You have NO BUFFER LEFT!
    - You have NO weak acid left!!

**Stop and check that it makes sense!**

Weak Acid + Strong Base  
Equivalence Point  
should be Basic  
Yes, 8.76 makes sense!

*Now you can do pH calculation!*

$$x = 5.77 \times 10^{-6} = [OH^-]$$

$$pOH = -\log(5.77 \times 10^{-6}) = 5.24$$

$$pH = 14 - pOH$$

$$pH = 14 - 5.24$$

$$**pH = 8.76**$$

# Titration Calculations

4

## AT ½ WAY POINT

- It took 37.5 ml of NaOH to get to eq.pt
  - So half way to eq. pt. would be 18.75 mL of NaOH

*@ ½ way point  $pH = pKa$*

$$pH = -\log(1.8 \times 10^{-5}) \quad \text{pH} = 4.74$$

**Why calculate pH at the ½ way point?** It is a nice point to plot on a graph to help get the curve. Also - when doing a titration, you can figure out the  $K_a$  by finding the pH at the halfway point.

# Titration Calculations

5

## AT THE END OF TITRATION

- Add 50mL of 0.10 M NaOH
  - Still no buffer anymore...used up all the weak acid.
  - BUT...you are past the equivalence point!
  - Now you have excess OH-

$$\frac{50\text{mL}}{1000\text{ mL}} \times \frac{0.10\text{ mol}}{1\text{ L}} = 5 \times 10^{-3}\text{ mol base}$$

*Do stoich to find how much left*

$\text{C}_2\text{H}_3\text{OH}$	$\text{OH}^- \leftrightarrow \text{H}_2\text{O}$	$\text{C}_2\text{H}_3\text{O}^-$
3.75mmol	5 mmol	0
0 mmol	$5 - 3.75 = 1.25\text{ mmol}$	3.75mmol

*Remember to use M in pH calculations!*

$$[\text{OH}^-] = \frac{1.25 \times 10^{-3}\text{ mol}}{(0.025\text{ L} + 0.050\text{ L})}$$

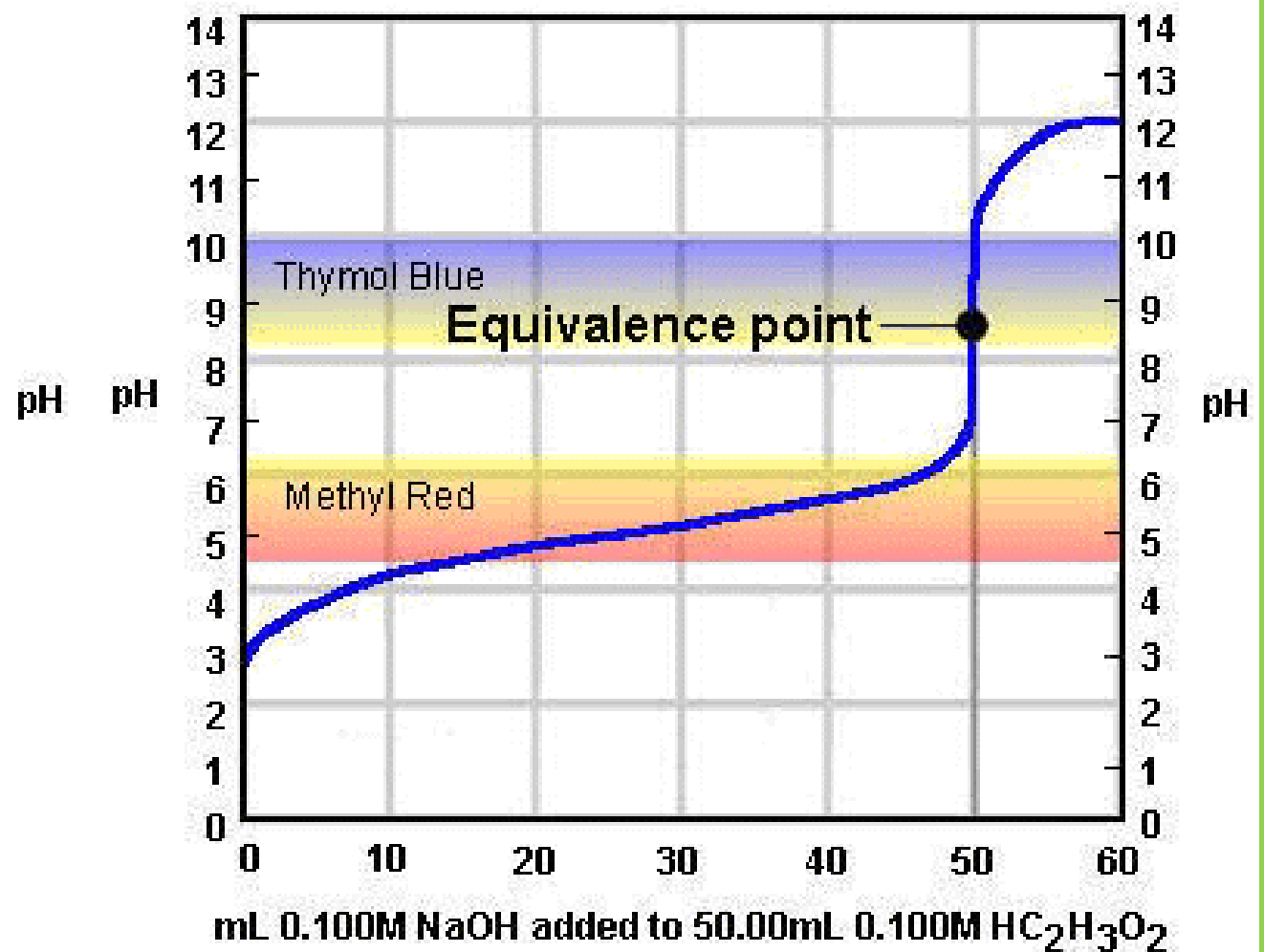
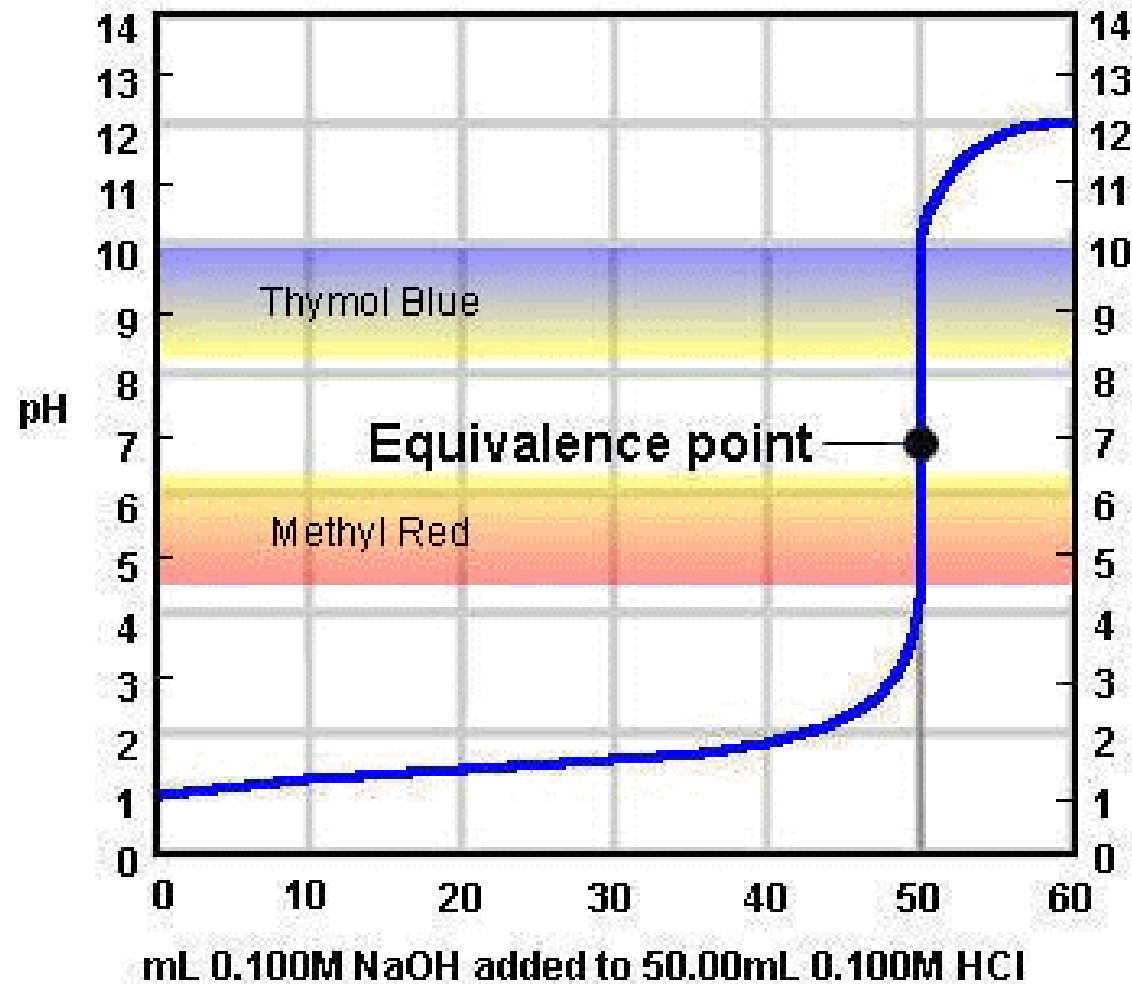
$$[\text{OH}^-] = 0.0167\text{ M}$$

$$p\text{OH} = -\log(0.0167) = 1.78$$

$$p\text{H} = 14 - p\text{OH} = 14 - 1.78$$

$$\text{pH} = 12.22$$

# Selection of Indicators


















# pH Indicators and Ranges

How do you know which indicator to pick for a reaction?

Pick the one that changes color in a pH range that is  $\pm 1$  from the pKa of the reaction you are doing.

Example: pKa 2

Choose something that changes in the 1 – 3 range.

pH Range	Color	Name
0.1-1.8		Crystal Violet
1.0-2.0		Cresol Red
1.2-2.8		Thymol Blue
2.7-4.0		2,4-Dinitrophenol
3.0-4.6		Bromophenol Blue
3.1-4.4		Methyl Orange
3.8-5.4		Bromocresol Green
4.2-6.3		Methyl Red
5.0-6.4		Eriochrome Black T
5.2-6.8		Bromocresol Purple
6.2-7.6		Bromothymol Blue
6.8-8.4		Phenol Red
6.8-8.6		m-Nitrophenol
8.3-10.0		Phenolphthalein
9.3-10.5		Thymolphthalein

# Some Acid-Base Indicators

Indicator	pH Range in which Color Change Occurs	Color Change as pH Increases
Crystal violet	0.0 - 1.6	yellow to blue
Thymol blue	1.2 - 2.8	red to yellow
Orange IV	1.4 - 2.8	red to yellow
Methyl orange	3.2 - 4.4	red to yellow
Bromcresol green	3.8 - 5.4	yellow to blue
Methyl red	4.8 - 6.2	red to yellow
Chlorophenol red	5.2 - 6.8	yellow to red
Bromthymol blue	6.0 - 7.6	yellow to blue
Phenol red	6.6 - 8.0	yellow to red
Neutral red	6.8 - 8.0	red to amber
Thymol blue	8.0 - 9.6	yellow to blue
Phenolphthalein	8.2 - 10.0	colourless to pink
Thymolphthalein	9.4 - 10.6	colourless to blue
Alizarin yellow	10.1 - 12.0	yellow to blue
Indigo carmine	11.4 - 13.0	blue to yellow

# YouTube Link to Presentation

**www.**