

Chemistry 12

Unit 4 – Acids, Bases and Salts

Unit Outline

1. Go over “The Arrhenius Theory of Acids and Bases” on p. 109 of SW.
In your notes, give the Arrhenius definition of and Acid, a Base and a Salt.
2. Do Ex. 1 on page 110 SW (check answers on page 279)
3. Go over steps for balancing formula equations for neutralization reactions on page 110 of SW. Do a couple of examples.
4. Do Ex. 2 a-f on page 110 SW (check answers on page 279)
5. Demonstration and notes on properties of acids and bases page 111 of SW
6. Do Ex 3 & 4 on p 112 of SW. (check answers on page 279)
7. Go over “Common Acids and Bases” on pages 112-114 of SW.
8. Do Worksheet 4-1 (On Chem 12 Web Page under Unit 4)
9. Do Tutorial 14 (On Chem 12 Web page) Make sure you do all the questions and check the answers with “Tutorial 14 Solutions”
10. Read pages 115-119 in SW
11. Do Ex 10a&b, 11, 12, 13, 14 & 15 p. 115-119
12. Read p. 119-120 in SW.
13. Do Ex 16-19 on page 121 of SW.
14. Have students print “Unit 4 Notes p 1-12” from Chem 12 Web page.
15. Do Demonstration on Conductivity of 1M HCl and 1M CH₃COOH. On “Unit 4 Notes” Go over the Section on “Strong and Weak Acids” p. 1-2
16. Take out your Data Booklet and look at the “RELATIVE STRENGTHS OF BRÖNSTED-LOWRY ACIDS AND BASES” table. (This can be called the “Acid Table”)
17. Carefully go through the section on “The Acid Table” from p 2-5 of Unit 4 Notes.
On page 4:
What is the [OH⁻] in 0.10 M Ba(OH)₂ ?
0.10 M **0.10M** **0.20 M**
Ba(OH)₂ → Ba²⁺ + 2OH⁻

And:
Find [OH⁻] in 0.10 M CaO
[O²⁻] = 0.10 M

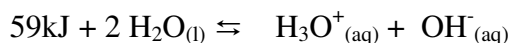
(0.10M) **0.20M**
O²⁻ + H₂O → 2 OH⁻

[OH⁻] = **0.20 M**
18. Go over Notes p. 6-7
19. Do Ex. 21-27 Pg.125-126 S.W.
20. Go over Section on **Acid-Base Equilibria & Relative Strengths of Acids & Bases** on pages 8-12 of Unit 4 Notes.

21. Do Ex. 38 - 46 p. 133 of SW.
22. Do Worksheet 4-2 on Chem 12 Web page.
23. Have students print Unit 4 Notes p. 13-31.
24. Go over p. 13-14 of Unit 4 Notes:

Questions:

Since reaction is endothermic:



At higher temps **products** are favoured and Kw is **high**er.

At lower temps **reactants** are favoured and Kw is **lower**.

Given: Kw at $60^\circ\text{C} = 9.55 \times 10^{-14}$

Calculate $[\text{H}_3\text{O}^+]$ & $[\text{OH}^-]$ at 60°C

In neutral water (assume water is neutral unless told otherwise) $[\text{H}_3\text{O}^+] = [\text{OH}^-]$

$$[\text{H}_3\text{O}^+][\text{OH}^-] = K_w = 9.55 \times 10^{-14}$$

Since $[\text{H}_3\text{O}^+] = [\text{OH}^-]$

$$[\text{H}_3\text{O}^+]^2 = 9.55 \times 10^{-14}$$

$$\text{So } [\text{H}_3\text{O}^+] = [\text{OH}^-] = \sqrt{9.55 \times 10^{-14}} = 3.09 \times 10^{-7}$$

25. Go over p. 15 in Unit 4 Notes.

Find $[\text{H}_3\text{O}^+]$ in 0.020 M $\text{Ba}(\text{OH})_2$

$$[\text{OH}^-] = 0.040 \text{ M}$$

$$[\text{H}_3\text{O}^+] = \frac{1.00 \times 10^{-14}}{0.040} = 2.5 \times 10^{-13} \text{ M}$$

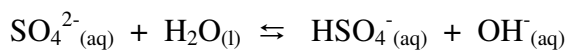
26. Go over "At Other Temps" on top of page 16
27. Rd. pg. 126-127 in SW and do Ex. 28-30 pg. 127 of SW.
28. Go over Section on pH on page 16-18 of Unit 4 Notes
Question on p. 18: Calculate pH of 12.0 M HCl

$$[\text{H}_3\text{O}^+] = 12.0 \text{ M}$$

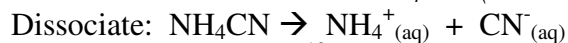
$$\text{pH} = -\log(12.0) = -1.079$$

29. Go over Sections on Converting pH \rightarrow $[\text{H}_3\text{O}^+]$ and Logarithmic Nature of pH on pages 18-19 on Unit 4 Notes
30. Carefully Go over the section on pOH and the relation between pH pOH Kw and pKw on pages 19-21 of Unit 4 notes.
31. Read p. 134-141 in SW.
32. Do ex. 49-53 + 55-57 (p. 139-141 S.W.)
33. Do Worksheet 4-3 pH and pOH Calculations on the Chem 12 Web page.

34. Go over the Section on **K_a & K_b for Weak acids and Bases** starting on p. 21 of Unit 4 Notes. Read to the top of page 24.
35. Read p. 148-151 in SW
36. Do ex. 79 & 81, p. 152 of SW
37. Go over p. 24-half way down p. 25 in Unit 4 Notes.
38. Do Ex. 77 & 80 on p. 152 SW
39. Go over half way down p. 25 to bottom of p. 26 in Unit 4 Notes
40. Do Ex. 76 & 78 on p. 152 of SW.
41. Go over “Now For Bases” from p. 26-31 in Unit 4 notes.
42. Rd. pg. 128 – 129 & 152-153 in SW
43. Do Ex. 84, 87, 88 & 89 on p. 153 of SW
44. Do Worksheet 4-4 K_a and K_b calculations for Chem. 12 Web page.
45. Have students print “Unit 4 Notes p. 32-53 from the Chem. 12 Web page.
46. Go over the Section on “**Hydrolysis**” from p. 32-top of p. 35 of Unit 4 Notes
The answer to the question on the top of p. 35 is:
*Write the **net-ionic equation** for the **hydrolysis** taking place in aqueous magnesium sulphate:*



47. Go over Section on “**Hydrolysis When BOTH Cation and Anion hydrolyze**” on p. 35 of Unit 4 Notes. The answer to the question near the bottom of p. 35 is:
Determine whether the salt NH_4CN (ammonium cyanide) is acidic, basic or neutral.

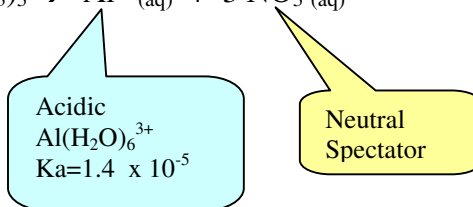
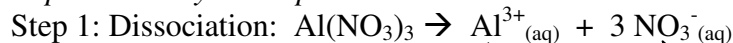


$$K_a \text{ of } \text{NH}_4^{+} = 5.6 \times 10^{-10}$$

$$K_b \text{ of } \text{CN}^{-} = \frac{1.0 \times 10^{-14}}{4.9 \times 10^{-10}} = 2.0 \times 10^{-5}$$

so since $K_b \text{ of } \text{CN}^{-} > K_a \text{ of } \text{NH}_4^{+}$ this solution is **basic**.

48. Go over “Hydrolysis of Amphotropic Anions” from p.35-36 in Unit 4 Notes.
49. Read p. 144 – 147 in SW & Do Ex. 69 (a-f) and Ex. 70 (a – j), 71, 72 & 73 on p. 148 SW.
50. Do Worksheet 4-5 (Hydrolysis) from Chem 12 Web page.
51. Do Experiment 20D—Hydrolysis (Print it from Chem 12 Web page.)
52. Go over “Putting it all Together—Finding the pH in a Salt Solution” from p. 36-37 in Unit 4 Notes. Do the Question on the top of p. 38.
The answer to the question on the top of p. 38:
Calculate the pH of a 0.24 M solution of the salt aluminum nitrate. Show all your steps. State any assumptions used.



So we treat this question like the pH of a Weak Acid ($\text{Al}(\text{H}_2\text{O})_6^{3+}$)

Step 2: Write HYDROLYSIS EQUATION (Don't forget that $\text{Al}(\text{H}_2\text{O})_6^{3+}$ undergoes ACID hydrolysis!)

And an ICE table underneath it:

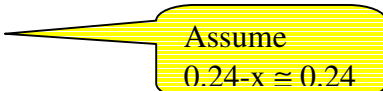
$$\text{Al}(\text{H}_2\text{O})_6^{3+}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{Al}(\text{H}_2\text{O})_5(\text{OH})^{2+}$$

[I]	0.24		0	0
[C]	- x		+ x	+ x
[E]	0.24 - x		x	x

Step 3: Write the K_a expression for the hydrolysis of $\text{Al}(\text{H}_2\text{O})_6^{3+}$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{Al}(\text{H}_2\text{O})_5(\text{OH})^{2+}]}{[\text{Al}(\text{H}_2\text{O})_6^{3+}]}$$

Step 4: Insert equilibrium concentration [E] values from the ICE table into the K_b expression. State any valid assumptions:

$$K_a = \frac{x^2}{(0.24 - x)}$$


Step 5: Calculate the value of x. Remember in the ICE table, that $x = [\text{H}_3\text{O}^+]$

$$K_a \cong \frac{x^2}{0.24}$$

$$1.4 \times 10^{-5} = \frac{x^2}{0.24}$$

$$x^2 = 0.24 (1.4 \times 10^{-5})$$

$$[\text{H}_3\text{O}^+] = x = \sqrt{0.24(1.4 \times 10^{-5})} = 1.83 \times 10^{-3}$$

Step 6: Calculate pH:

$$\text{pH} = -\log (1.83 \times 10^{-3}) = \mathbf{2.74}$$

53. Do Demonstration in obtaining the pH's of various oxide solutions. Use Table on p. 38. Go over the Section "Metal, Non-metal and Metalloid Oxides (also called Anhydrides)" p. 38 to half way down p. 39 of Unit 4 Notes. The important blanks are filled in on the next page:

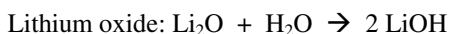
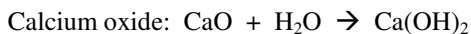
Compound	Metal or Non-metal Oxide	Colour in Universal Indicator	Approximate pH
Aqueous MgO	Metal oxide		~10
Aqueous CaO	Metal oxide		~9
Aqueous ZnO	Metal oxide		~8
Aqueous CO ₂	Non-metal oxide		~5
Aqueous NO ₂	Non-metal oxide		~3
Aqueous SO ₂	Non-metal oxide		~3

Conclusions:

Metal oxides act as bases in aqueous solution.

Non-Metal oxides act as acids in aqueous solution.

Write a balanced formula equation for the overall reactions of the following oxides with water:



54. Quickly go over the material from the middle of p. 39 to the middle of p. 42 in Unit 4 Notes. Students will be responsible for this.

55. Go over “Review of Titrations” from p. 42 to p. 44 in Unit 4 Notes.

Here is the answer to the question on the top of p. 44:

0.200 M NaOH is used to titrated 3 separate 50.0 mL samples of a solution of H₂SO₄ of unknown concentration.

The NaOH is in the burette. Use the following data table to calculate the [H₂SO₄] in the original H₂SO₄ solution. Show all of your steps clearly, including the balanced formula equation for the reaction.

	Trial 1	Trial 2	Trial 3
Initial Burette Reading (mL)	0.00	9.02	17.95
Final Burette Reading (mL)	9.02	17.95	26.89
Volume of NaOH used (mL)	9.02	8.93	8.94

Discard the 9.02 volume value and take the average of 8.93 and 8.94, which is 8.935 mL

Calculate the moles of NaOH: $0.008935 \text{ L} \times 0.200\text{M} = 0.001787 \text{ mol NaOH}$

Write the Balanced Equation: $2 \text{NaOH} + \text{H}_2\text{SO}_4 \rightarrow 2 \text{H}_2\text{O} + \text{Na}_2\text{SO}_4$

Calculate the moles of H₂SO₄:

$$0.001787 \text{ mol NaOH} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}} = 0.0008935 \text{ mol H}_2\text{SO}_4$$

Calculate the [H₂SO₄]:

$$[\text{H}_2\text{SO}_4] = \frac{0.0008935 \text{ mol H}_2\text{SO}_4}{0.050 \text{ L H}_2\text{SO}_4} = \mathbf{0.0179 \text{ M}}$$

56. Read “Experimental Note” on the bottom of page 157 of SW.
 57. Do Ex. 94 – 97 on p. 158 of SW.

NOTE: If you check your answers on p. 291 of SW, you’ll notice, he uses “mmol’s” (millimoles) in his work.

See p. 154 – 155 of doing these using mmols. You don’t have to use mmols. You can if you like. Just be careful with your units.

58. Do Worksheet 4-6 from the Chem. 12 Web page.
 59. Go over the section on Indicators from p. 44 to half way down p. 46 on Unit 4 Notes.
 60. Do the question in the middle of p. 46 of Unit 4 Notes.

Here are the answers to check:

*Question: When a drop of 0.1M HCl is added to the indicator bromcresol green, the colour is yellow.
 When a drop of 0.10M NaOH is added to the indicator, the colour is blue.*

- a. What colour is the acid form of bromcresol green (HInd)? yellow
 b. What colour is the base form of bromcresol green (Ind⁻)? blue
 - What would the colour be if [HInd] = [Ind⁻] for bromcresol green? Green
 61. Go over the section on “Transition Point” from the bottom of p. 46 to the top of p. 49. Do all the questions within that section. The answers are as follows:

So the Acid Form of methyl violet (HInd) is YELLOW and the Base Form of methyl violet (Ind⁻) is BLUE.

The colour at the TRANSITION POINT of

Methyl violet would be green

The colour at the TRANSITION POINT of

Bromcresol green would be green

The colour at the TRANSITION POINT of

Indigo carmine would be green

pH	Indicator	
	Thymol Blue	Orange IV
0.8	red	red
2.0	orange	orange
3.5	yellow	yellow

62. Go over the section on “**Finding the Ka of an Indicator**” from p. 49 to p. 51 of Unit 4 Notes. Do all the questions in that section. The answers are here to check yours with (starting on the next page...)

Find the K_a of Alizarin Yellow:

$$pK_a = \frac{10.1 + 12.0}{2} = 11.05$$

$K_a = \text{antilog}(-pK_a) = \text{antilog}(-11.05) = 9 \times 10^{-12}$ (1 SD because pH's on Indicator Table are only to 1 decimal place.)

This is because Thymol Blue is a diprotic acid. Each time it loses a proton, it goes through a color change.

We can call Thymol Blue (Tb) a weak acid H_2Tb

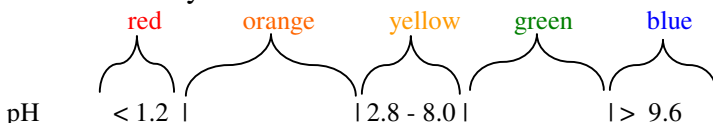
The equilibrium equation for the **first ionization** is: $H_2Tb + H_2O \rightleftharpoons H_3O^+ + HTb^-$
Using the table above, fill in the colours: red yellow

The equilibrium equation for the **second ionization** is: $HTb^- + H_2O \rightleftharpoons H_3O^+ + Tb^{2-}$
Using the table below, fill in the colours: yellow blue

pH	Form(s) which predominate(s) (H_2Tb , HTb^- or Tb^{2-})	Approximate Colour
1.0	H_2Tb	red
2.0	H_2Tb & HTb^- are equal	orange
3.0	HTb^-	yellow
7.0	HTb^-	yellow
8.8	HTb^- & Tb^{2-} are equal	green
10.0	Tb^{2-}	blue

Also, fill in the colours on the following diagram:

Colours of Thymol Blue:



Find the pH's and the colours of the given indicators in the following solutions (assume temp. = 25°C):

Solution	pH	Colour in Thymol Blue	Colour in Methyl Red	Colour in Alizarin Yellow
0.2 M HCl	0.7	red	red	yellow
0.01 M HCl	2.0	orange	red	yellow
0.0005 M HCl	3.3	yellow	red	yellow
Pure water	7.0	yellow	yellow	yellow
0.0001 M NaOH	10.0	blue	yellow	yellow
0.2 M NaOH	13.3	blue	yellow	red

For example, a solution displays the following colours in the indicators shown. See if you can narrow the pH down to a range:

Indicator	Colour of Solution	Approximate pH Range
Bromthymol blue	Blue	> 7.6
Thymol blue	Yellow	2.8 – 8.0
Phenolphthalein	Colourless	< 8.2
Approximate pH range of the solution using all information:		7.6 – 8.0

Let's try another one:

For example, a solution displays the following colours in the indicators shown. See if you can narrow the pH down to a range:

Indicator	Colour of Solution	Approximate pH Range
Orange IV	Yellow	> 2.8
Methyl red	Red	< 4.8
Methyl Orange	Red	< 3.2
Approximate pH range of the solution using all information:		2.8 – 3.2

For example, a solution displays the following colours in the indicators shown. See if you can narrow the pH down to a range:

Indicator	Colour of Solution	Approximate pH Range
Methyl Orange	Yellow	> 4.4
Alizarin Yellow	Yellow	< 10.1
Thymol Blue	Green	8.0 – 9.6
Approximate pH range of the solution using all information:		8.0 – 9.6

63. Read p. 159 – 162 in SW.

64. Do Ex. 108 – 112 and 114 – 120 on p. 162 – 163 of SW

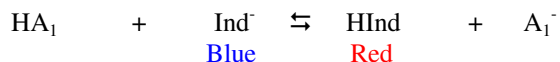
65. Go over the section “Using Indicators to Rank Weak Acids in Order of Strengths” from p. 52 – 53 of Unit 4 Notes. Answer all the questions in that section. The answers are here (starting on the next page...)

A few drops of this indicator (a mixture of HInd and Ind⁻) is added to a weak acid called HA₁ and the colour is **blue**.

Which is the stronger acid, HA₁ or HInd?

To find out, we write an equilibrium equation (NOT with H₂O this time!). For reactants, we use the weak acid HA₁ and the **base** form of the indicator, Ind⁻. (two acids are not written on the same side of equilibrium equations!)

I'm sure you can fill in the two products: Write the colours of Ind⁻ and HInd right underneath each one.



Since the colour of the indicator was **blue**, it means that the form of the indicator (*HInd or Ind⁻*) **Ind⁻** is predominating (favoured by the equilibrium). So the (*reactants/products*) **reactants** of the

equation above are favoured, meaning (*HA₁/HInd*) **HA₁** is the Weaker acid or (*HA₁/HInd*) **HInd** is the Stronger Acid.

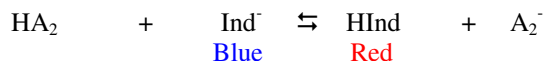
Now let's look at another experiment involving the same indicator and a different weak acid HA₂.

A few drops of this indicator (a mixture of HInd and Ind⁻) is added to a weak acid called HA₂ and the colour is **red**.

Which is the stronger acid, HA₂ or HInd?

To find out, we write an equilibrium equation (NOT with H₂O this time!). For reactants, we use the weak acid HA₂ and the **base** form of the indicator, Ind⁻. (two acids are not written on the same side of equilibrium equations!)

I'm sure you can fill in the two products: Write the colours of Ind⁻ and HInd right underneath each one.



Since the colour of the indicator was **red**, it means that the form of the indicator (*HInd or Ind⁻*) **HInd** is predominating (favoured by the equilibrium). So the (*reactants/products*) **products** of the

equation above are favoured, meaning (*HA₂/HInd*) **HInd** is the Weaker acid or (*HA₂/HInd*) **HA₂** is the Stronger Acid.

So, to summarize the results of both experiments: Experiment 1: **HInd > HA₁**

Experiment 2: **HA₂ > HInd**

So, in comparing strengths of HA₁ and HA₂, we can say that: **HA₂ > HA₁**

Now, make a little mini acid table with the acids on the left, a \rightleftharpoons in the middle and H⁺ + conj. base on the right.

Put the acids in order of strongest → weakest.

	Acid		Base	
↑ stronger	HA ₂	\rightleftharpoons	H ⁺ + A ₂ ⁻	↓ stronger
	HInd	\rightleftharpoons	H ⁺ + Ind ⁻	
	HA ₁	\rightleftharpoons	H ⁺ + A ₁ ⁻	

66. Do Worksheet 4-7 from the Chem. 12 Web page.
 67. Have students print Unit 4 Notes p. 54 –77 from Chem. 12 Web page.
 68. Go over Section on “Practical Aspects of Titration” p. 54 to part way down p. 55 of Unit 4 Notes. Do all the questions in that section. The answers to those questions are here for you to check:

Eg.) 40.48 g of potassium hydrogen phthalate ($\text{KHC}_8\text{H}_4\text{O}_4$) is weighed out and dissolved in enough distilled water to make 1.000 L of solution. Find the $[\text{KHC}_8\text{H}_4\text{O}_4]$.
 (HINT: Use $\text{g} \rightarrow \text{moles} \rightarrow \text{M}$)

$$40.48 \text{ g} \times \frac{1 \text{ mol}}{204.1 \text{ g}} = 0.1983 \text{ mol}$$

$$[\text{KHC}_8\text{H}_4\text{O}_4] = \frac{0.1983 \text{ mol}}{1.000 \text{ L}} = \mathbf{0.1983 \text{ M}}$$

It takes 4.02 mL of 0.200 M $\text{KHC}_8\text{H}_4\text{O}_4$ to titrate 10.00 mL of a solution of NaOH. Find the $[\text{NaOH}]$

The balanced equation for the reaction is : $\text{KHC}_8\text{H}_4\text{O}_4 + \text{NaOH} \rightarrow \text{H}_2\text{O} + \text{KNaC}_8\text{H}_4\text{O}_4$

moles of $\text{KHC}_8\text{H}_4\text{O}_4 = 0.200 \text{ M} \times 0.00402 \text{ L} = 0.000804 \text{ mol KHC}_8\text{H}_4\text{O}_4$

moles of NaOH = $0.000804 \text{ mol KHC}_8\text{H}_4\text{O}_4 \times \frac{1 \text{ mol KHC}_8\text{H}_4\text{O}_4}{1 \text{ mol NaOH}} = 0.000804 \text{ mol NaOH}$

$$[\text{NaOH}] = \frac{0.000804 \text{ mol NaOH}}{0.01000 \text{ L NaOH}} = \mathbf{0.0804 \text{ M}}$$

So the $[\text{NaOH}] = \mathbf{0.0804 \text{ M}}$

Eg.) It takes 28.54 mL of standardized 0.0804 M NaOH to titrate a 25.00 mL sample of an H_2SO_4 solution.

The balanced equation for this neutralization reaction is: $2\text{NaOH} + \text{H}_2\text{SO}_4 \rightarrow 2\text{H}_2\text{O} + \text{Na}_2\text{SO}_4$

Calculate the $[\text{H}_2\text{SO}_4]$.

moles of NaOH = $0.0804 \text{ M} \times 0.02854 \text{ L} = 0.0022946 \text{ mol NaOH}$

moles of $\text{H}_2\text{SO}_4 = 0.0022946 \text{ mol NaOH} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}} = 0.0011473 \text{ mol H}_2\text{SO}_4$

$$[\text{H}_2\text{SO}_4] = \frac{0.0011473 \text{ mol H}_2\text{SO}_4}{0.02500 \text{ L H}_2\text{SO}_4} = \mathbf{0.0459 \text{ M}}$$

Answer $[\text{H}_2\text{SO}_4] = \mathbf{0.0459 \text{ M}}$

69. Carefully read p. 164 – 165 in SW.
 70. Do Ex. 121 – 123 on p. 165 of SW.
 71. Go over the section “**Finding the pH of Mixtures of Acids and Bases**” from pag 55 to 1/3 of the way down p. 66 in Unit 4 Notes. Do the question near the top of page 66. Here is the answer:

40.00 mL of 0.100 M NaOH is mixed with 25.00 mL of 0.100 M HCl. Calculate the pH of the solution resulting. Show all of your steps. Express your answer in the correct # of SD's as justified by the data.

Solution: Balanced equation: $\text{NaOH} + \text{HCl} \rightarrow \text{H}_2\text{O} + \text{NaCl}$

Initial moles of NaOH : $0.100 \text{ M} \times 0.04000 \text{ L} = 0.00400 \text{ mol NaOH}$ (3 SD's like the 0.100 M) (5 dec. places)

Initial moles HCl : $0.100 \text{ M} \times 0.02500 \text{ L} = 0.00250 \text{ mol HCl}$ (3 SD's like the 0.100 M) (5 dec. places)

Excess moles: $\text{NaOH} = 0.00150 \text{ mol NaOH}$ (5 dec. places) (3 SD's)

Volume of final mixture: $40.00 \text{ mL} + 25.00 \text{ mL} = 65.00 \text{ mL} = 0.06500 \text{ L}$
 (2 dec. pl.) (2 dec. pl.) (2 dec. pl.) (4 SD's)
 (4 SD's)

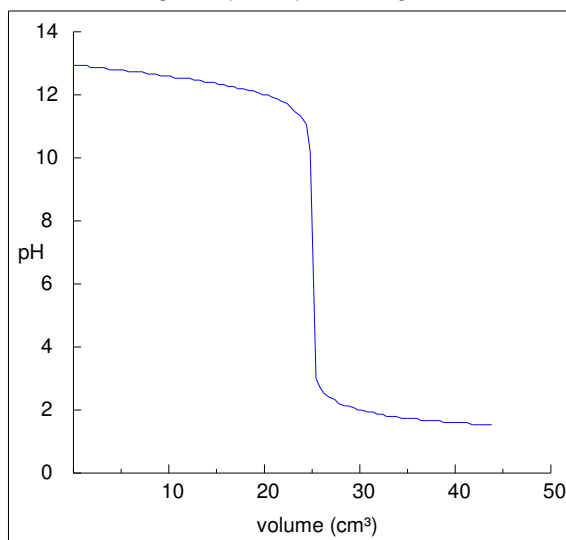
$[\text{OH}^-] = [\text{NaOH}] \text{ in the final mixture} = \frac{0.00150 \text{ mol (3 SD's)}}{0.06500 \text{ L (4 SD's)}} = \underline{0.023077 \text{ M}}$

pOH = 1.63682

pH = 12.363 (3 SD's)

72. Go over the example for “Type 2” on the rest of page 66 of Unit 4 Notes.
 73. Do Ex. 58, 59 and 60 on p. 143 of SW
 74. Go over the section on “Titration Curves” from p. 57 – 59 of Unit 4 Notes.
 75. Do the calculations and fill in the table on the bottom of page 59. Students can do these in groups with each student assigned a few calculations.
 76. Follow the instructions on page 60 of Unit 4 Notes and make a graph of this Strong Acid-Strong Base Titration Curve.
 77. In order to check your results, go to the site:
<http://wwwchem.uwimona.edu.jm:1104/software/titr.html> (you can also get to it from the Chem 12 web site. (called “Titration Simulation Site”)
 When it comes up, scroll down and click on “MS Windows Version...”
 Click “OPEN” when the file download menu comes up.
 Go to “Titration” → “Reactants” → “Strong Acid with Strong Base
 Leave “Molarity of Acid and Molarity of Base” as 0.1 M. Click OK
 Click on magnetic stirring bar to get it moving.
 Click on the stopcock to gradually add the base to the acid. The graph will draw itself. It should look similar to the graph you got following the instructions on page 60 of Unit 4 Notes.
 78. Carefully Go over the section on “**Weak Acid—Strong Base Titration Curves**” from page 60-64 in Unit 4 Notes. Make sure you familiarize yourself with the characteristics of a WA-SB titration curve and how it is different from a SA-SB curve.

79. Go over the Section on **“Strong Acid—Weak Base Titration Curves”** on page 65 of Unit 4 Notes.
80. To find out what a SA-WB Titration Curve looks like, go back to the “Titration Simulation Site” and this time, under “Titrations” → “Reactants”, scroll to “Weak Acid with Strong Base” and Click “OK” for both boxes that come up. Keep clicking on the stopcock to add the strong base to the weak acid. You might want to print a copy of the graph. NOTICE that the pH at Equivalence Point is a bit BELOW 7.
81. Go over the “Summary” on the top of page 66 and the Section on “Indicators for Titrations” from p. 66 – 67
82. Do the question on the top of page 68.
The answer should look something like this:

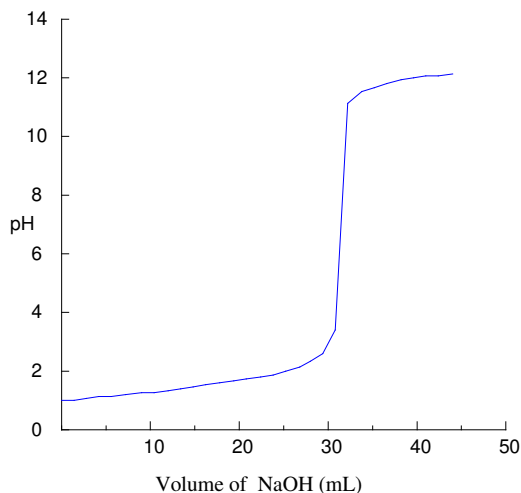
Titration of strong base (25 cm³) with strong acid.

Name an indicator which would be suitable for this titration. Bromthymol Blue

As you pass through the Equivalence (Stoichiometric) point in this titration, the colour of your indicator would change from blue to green (to yellow)

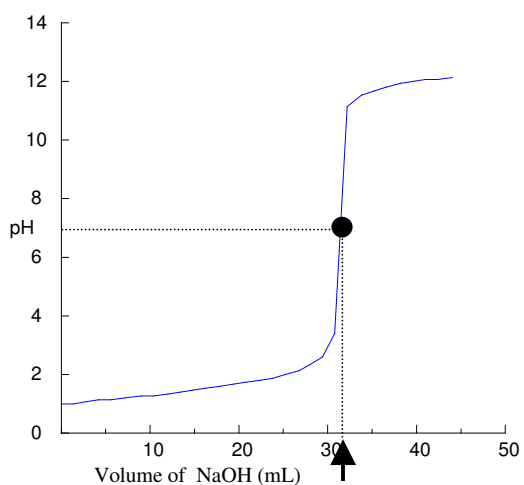
83. Go over the section on **“Using Titration Curves in Questions and Calculations”** from page 68 to the bottom of page 70.
Do all the questions that are in that section.
(Answers to questions start on the next page...)

When titrating a 25.0 mL sample of 0.10 M HCl with a solution of NaOH, the following titration curve was obtained. Calculate the [NaOH] in the burette:



We know that this is a SA-SB titration, so at the EP, pH = 7.

Also, the EP is always in the center of the “almost vertical” region. We mark the EP and draw a straight line down to see where it hits the “Volume of Base” axis. This will give us the Volume of NaOH needed to reach the equivalence point:



We see that the Volume of NaOH needed to reach the Equivalence Point is approximately 32 mL.

Given this and the information at the beginning of the question, calculate the [NaOH] in the burette:

Calculation:

$$\text{moles of HCl} = 0.10 \text{ M} \times 0.0250 \text{ L} = 0.0025 \text{ mol HCl}$$

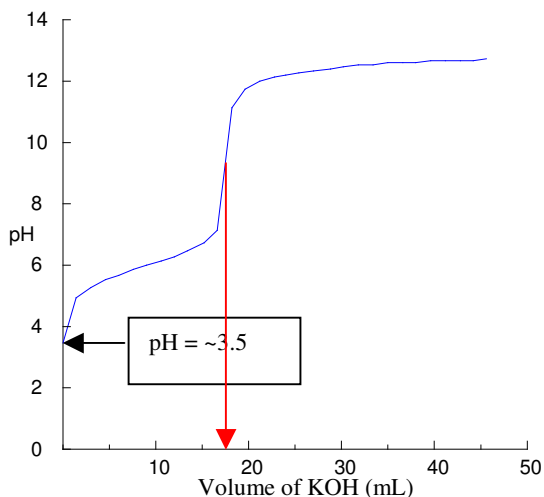
$$\text{moles of NaOH} = \text{moles of HCl (coefficient ratio in equation)} = 0.0025 \text{ moles NaOH}$$

$$[\text{NaOH}] = \frac{0.0025 \text{ mol NaOH}}{0.032 \text{ L NaOH}} = \mathbf{0.078 \text{ M}}$$

$$[\text{NaOH}] = \mathbf{0.078 \text{ M}}$$

Question:

The following titration curve results from titrating 25.0 mL of a 0.10 M Weak Acid HA with a Strong Base KOH:



a.) Use this graph to estimate the K_a of the acid HA.

pH (of 0.10M acid with no base) = 3.5

$$[H_3O^+] = 3.16 \times 10^{-4} \text{ M}$$

$$K_a = \frac{[H_3O^+]^2}{(C_o - [H_3O^+])} = \frac{(3.16 \times 10^{-4})^2}{(0.10 - 3.16 \times 10^{-4})}$$

$$K_a = 1.0 \times 10^{-6}$$

b.) Use this graph to calculate the [KOH].

Volume of KOH required to reach EP ~ 17 mL

$$\text{moles HA} = 0.10 \text{ M} \times 0.025 \text{ L} = 0.0025 \text{ mol HA}$$

$$\begin{aligned} \text{moles KOH} &= \text{moles HA (1:1 coefficient ratio)} \\ &= 0.0025 \text{ mol KOH} \end{aligned}$$

$$[\text{KOH}] = \frac{0.0025 \text{ mol}}{0.017 \text{ L}} = 0.15 \text{ M}$$

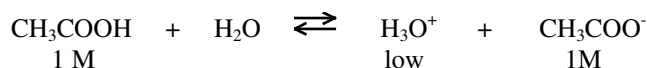
84. Go over the section on “**More Practical Things about Titrations**” on the bottom of page 70 of Unit 4 Notes.

85. Go through the section on “**Buffer Solutions**” from p. 71-75 of Unit 4 Notes. Do all the questions that are in that section as you are going through it.

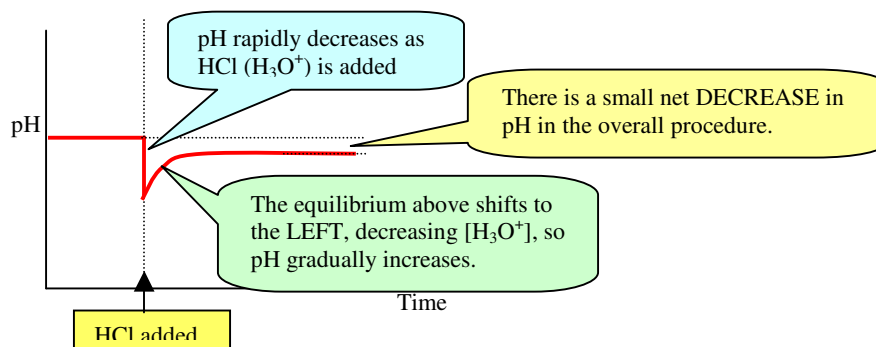
The answers to the questions are as follows:

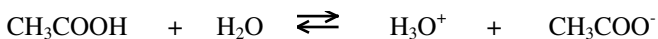
Consider the buffer solution made up of 1M CH_3COOH and 1M NaCH_3COO .

The equation representing the equilibrium present in this buffer solution is (just the same as the ionization of the weak acid, CH_3COOH):

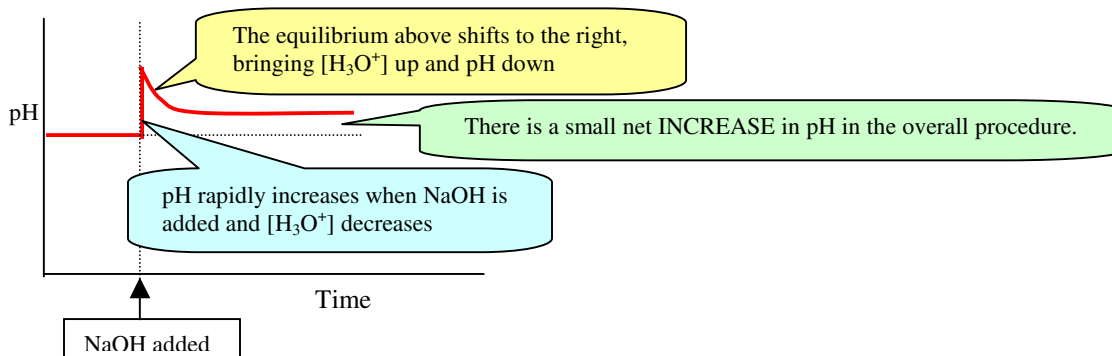


Draw a graph of **pH vs. Time** when a small amount of HCl is added to the buffer above (Explain each part)





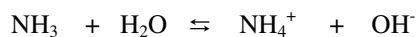
Draw a graph of **pH vs. Time** when a small amount of NaOH is added to the buffer above (Explain each part)



Basic Buffers work using the same principles. Work through the example by filling in the blanks.....

Eg.) A buffer solution is prepared using 1M NH_3 and 1M NH_4Cl (WBSCA)

a) Write the **equilibrium equation** describing this buffer.

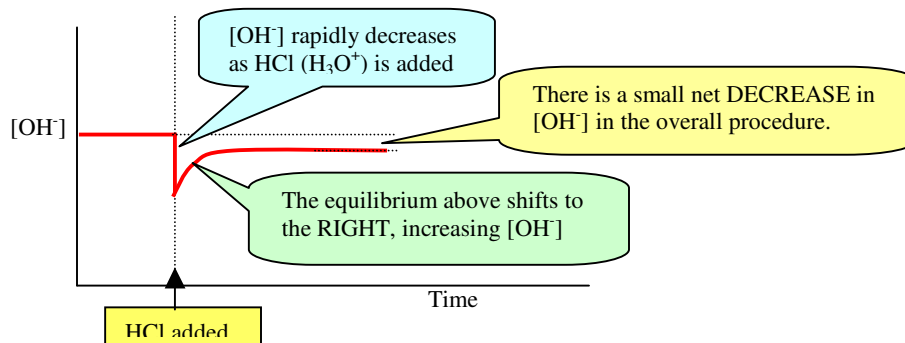


b) When a small amount of HCl (SA) is added, the $[\text{OH}^-]$ quickly **decreases** (the pH goes **down**)

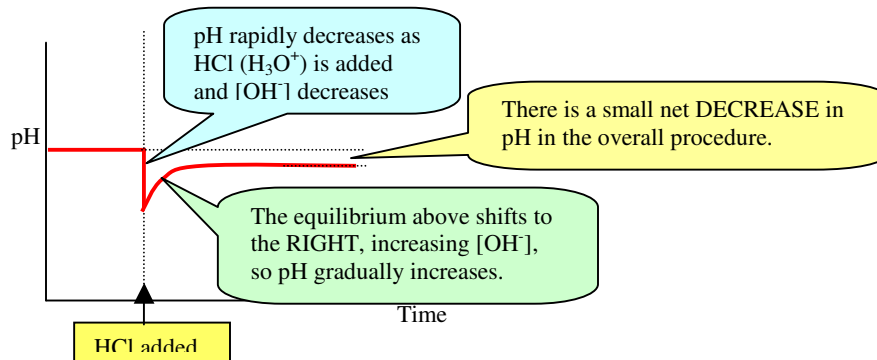
c) As a result, the equilibrium shifts to the **right**, and the $[\text{OH}^-]$ gradually **increases**. (the pH goes back **up**)

d) So, as a result of adding HCl, there was a small **net decrease** in the $[\text{OH}^-]$ (a small **net decrease** in pH)

e) Draw a graph of $[\text{OH}^-]$ vs. Time to illustrate what happened in b \rightarrow d. Label each part.

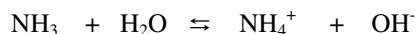


f) Draw a graph of pH vs. Time to illustrate what happened in b → d.



Eg.) A buffer solution is prepared using 1M NH₃ and 1M NH₄Cl (WBSCA)

a) Write the **equilibrium equation** describing this buffer.

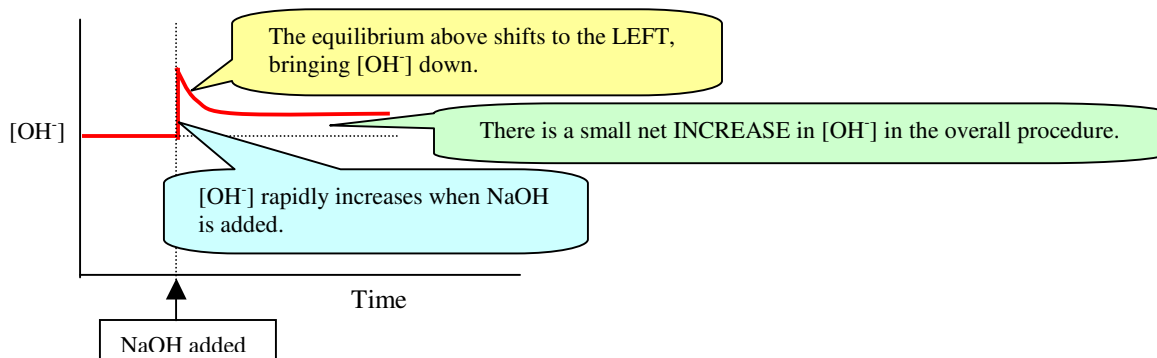


b) When a small amount of NaOH (SB) is added, the [OH⁻] quickly **increases** (the pH goes **up**)

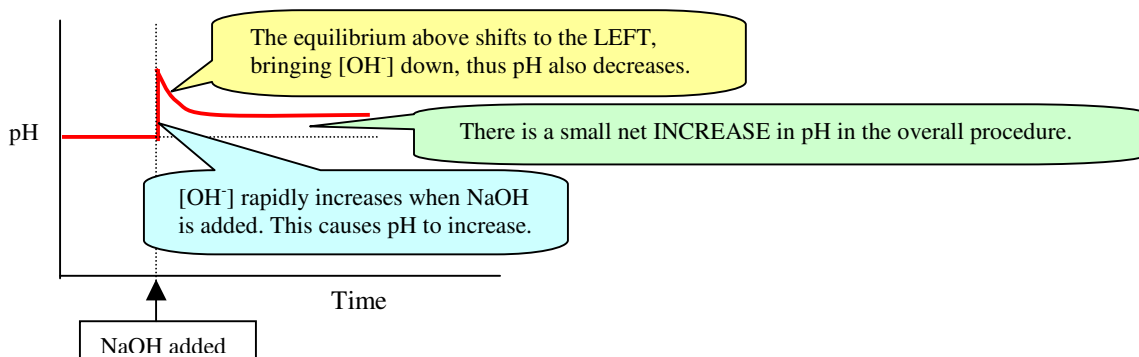
c) As a result, the equilibrium shifts to the **left**, and the [OH⁻] gradually **decreases**. (the pH goes back **down**)

d) So, as a result of adding NaOH, there was a small **net increase** in the [OH⁻] (a small **net increase** in pH)

e) Draw a graph of [OH⁻] vs. Time to illustrate what happened in b → d. Label each part.



f) Draw a graph of pH vs. Time to illustrate what happened in b \rightarrow d.



So, in summary, this buffer **minimizes (or resists) changes in pH** when a small amount of acid or base is added to it!

86. If time permits, do Experiment 20F—Buffers Parts 2 and 3 as a demonstration.
87. Go over the Sections on “Limitations of Buffers” and “Biological Buffer Systems” on p. 76-77 of Unit 4 Notes.
88. Have students do the “Provincial Questions for Unit 4” review sheet. This is not on the website.