Acid & Base Math Made Easy

In acid/base chemistry, there are a variety of different questions that may be asked. It may be helpful to think of these questions as a puzzle in which you are trying to solve for the missing piece. Similarly to a puzzle, sometimes it's difficult and overwhelming to know where to begin. So here's a little outline to help! But first, remember to succeed in this unit it is **incredibly important** to be able to **identify** strong versus weak acids and bases.

**Forewarning, this does not cover every explicit math problem that may be asked, but some examples of common scenarios and what to do. Please make sure to use the homeworks, PLQs, recitation worksheets, lecture activities, etc. when preparing for Exam 3. 😊

Scenario 1: Strong Acid <u>OR</u> Strong Base

If a question pertains to finding the concentration or pH/pOH of a strong acid (SA) OR strong base (SB) whatever concentration you begin with COMPLETELY goes to products. This is because strong electrolytes **completely dissociate**.

For example, say a question asks about the pH of a 0.8 M NaOH solution:

$$\begin{array}{ccccc} NaOH\ (aq) & \rightarrow & Na^{+}\ (aq) & + & OH^{-}\ (aq) \\ 0.8\text{ M} & & \text{(Negl.)} & & 0.8\text{ M} \end{array}$$

Thus, the concentration of hydroxide is 0.8 M. The pH can be found by calculating the pOH and subtracting it by 14 (assuming this is at 25°C).

$$pOH = -\log(0.8) = 0.0969$$

$pH = 14 - 0.0969 =$ **13.90** (This pH makes sense as it is basic)

Look out for **Group II hydroxides. These produce **twice** as much hydroxide. It is always important to draw the dissociation reaction to ensure the concentrations are stoichiometrically correct.

Practice below by finding the pH of 0.75 M $Ba(OH)_2$ solution at 25°C:

Scenario 2: Strong Acid <u>AND</u> Strong Base

If a question pertains to finding the concentration or pH/pOH of a SA mixed with a SB this is a **titration problem**. The acid and base may have **different volumes** so we will use an **ICF table** which allows us to compare the MOLES of each. Remember ICF tables help us determine the **LIMITING REACTANT** which will be **completely consumed**. After completing the ICF table this will leave you with **either H^+ OR OH^-** remaining OR, if at the *equivalence point,* **neither H^+ OR OH^-** (then the pH = 7!). We will focus on the **former scenario** here.

For example, say a question asks about the pH of a solution that has been titrated with 100 mL of 0.05 M HCl into 200 mL solution of 0.01 M NaOH:

1. Find H^+ and OH^- moles. (Remember for Group II hydroxides it will be double)

HCl (*aq*)	→	H^+ (*aq*)	+	Cl^- (*aq*)
0.05 M		0.05 M		(Negl.)

H^+ from HCl: 0.1 L x 0.05 mol/L = **0.005 mol H^+**

$NaOH$ (*aq*)	→	Na^+ (*aq*)	+	OH^- (*aq*)
0.01 M		(Negl.)		0.01 M

OH^- from NaOH: 0.2 L x 0.01 mol/L = **0.002 mol OH^-**

NaOH is the limiting reactant (less moles)

2. Write the reaction and fill in the ICF table

	OH^- (*aq*)	+	H^+ (*aq*)	→	H_2O (ℓ)
I	0.002 mol		0.005 mol		\
C	−0.002 mol		−0.002 mol		\
F	0		**0.003 mol**		\

Only H^+ remains. To find the pH we need to convert back to concentration with the **NEW volume**.

$[H^+]$: 0.003 mol / (0.1 L + 0.2 L) = **0.01 M**

pH = -log(0.01) = **2** (This pH makes sense as it is acidic)

Practice below by finding the pH of a solution that has been titrated with 250 mL of 0.5 M HBr into 250 mL solution of 2 M KOH:

Scenario 3: Weak Acid **OR** Weak Base

If a question pertains to finding the concentration or pH/pOH of a weak acid (WA) OR weak base (WB) an ICE table must be used. This is because weak electrolytes **partially dissociate** thus the hydrolysis reaction is an **equilibrium process**.

For example, say a question asks about the pH of a 1.75 M CN^- solution (K_b = 1.6 x 10^{-5} at 25°C):

1. Write the hydrolysis and fill in the ICE table

	CN^- (*aq*)	+	H_2O (ℓ)	⇋	OH^- (*aq*)	+	HCN (*aq*)
I	1.75 M		\		0		0
C	−x		\		+x		+x
E	1.75−x		\		x		x

2. Solve for x using the K_b (use assumption method because $K_b \leq 10^{-5}$)

$$1.6 \times 10^{-5} = x^2 / 1.75$$

$$x = 0.00529 \text{ M} = [OH^-]$$

3. Calculate the pOH, then subtract from 14 to find the pH

$$pOH = -\log(0.00529) = 2.276$$

pH = 14 – 2.276 = **11.72** (This pH makes sense as it is basic)

Practice below by finding the pH of a 0.25 M HCN solution at 25°C (Hint: Use the K_b above to find the needed K_a):

Scenario 4: Weak Acid <u>AND</u> Weak Base

If a question pertains to finding the pH/pOH of a WA AND WB this is a **buffer solution** and the **Henderson-Hasselbalch equation (H-H Equation)** will be used. This is because a buffer is a mixture of **weak acid** and **weak base** (Conjugate Pairs). So, no ICE or ICF table is needed!

The H-H equation is available on your datasheet, so you do not need to memorize it. It takes the form of:

$$pH = pK_a + \log([Base]/[Acid])$$

For example, say a question asks about the pH of a buffer solution that consists of 0.25 M acetic acid and 0.15 M acetate ion (K_a = 1.75 x 10^{-5} at 25°C):

1. Calculate the pK_a from the K_a

$$pK_a = -\log(K_a) = -\log(1.75 \times 10^{-5}) = \mathbf{4.76}$$

2. Plug and chug values in the H-H equation

$$pH = 4.76 + \log([0.15\ M]/[0.25\ M]) = 4.53$$

Practice below by finding the pH of a buffer solution that consists of 1.75 M HCN and 1.05 M CN^- solution at 25°C (Hint: Use the K_b from Scenario 3 to find the needed K_a):

Scenario 5: Weak Acid/Base <u>**AND**</u> Strong Base/Acid

If a question pertains to finding the concentration or pH/pOH of a WA/WB mixed with a SB/SA this is a **titration problem**. The acid and base may have **different volumes** so we will use an **ICF table** which allows us to compare the MOLES of each. After completing the ICF table this will leave you with 1 of 3 different scenarios:

I. Only the **Weak Base** remaining (Leads to SCENARIO 3)

For example, say a question asks about the pH of a 100 mL solution of 1 M NH_4^+ that is titrated with 200 mL of 0.5 M NaOH (K_b = 1.8 x 10^{-5} of NH_3):

1. Find mols of each reactant to determine the limiting reagent

NaOH (*aq*)	→	Na^+ (*aq*)	+	OH^- (*aq*)
0.5 M		(Negl.)		0.5 M

OH^- from NaOH: 0.2 L x 0.5 mol/L = **0.1 mol OH^-**

Weak acids do not completely dissociate, so we will calculate the mols only for NH_4^+: 0.1 L x 1.0 mol/L = **0.1 mol NH_4^+**

NaOH and NH_4^+ are the limiting reactant (equal moles)

2. Write the reaction and fill in the ICF table

	OH^- (*aq*)	+	NH_4^+ (*aq*)	→	NH_3 (*aq*)	+	H_2O (ℓ)
I	0.1 mol		0.1 mol		0		\
C	−0.1 mol		−0.1 mol		+0.1 mol		\
F	0		0		0.1 mol		\

Only weak base (NH_3) remains, so we will follow the steps in **SCENARIO 3** to determine the pH of this solution. But first we need the concentration with the NEW VOLUME because we will be using an **ICE table**

$[NH_3]$: 0.1 mol / (0.1 L + 0.2 L) = **0.33 M**

3. Follow the steps in **SCENARIO 3**

	NH_3 (*aq*)	+	H_2O (ℓ)	⇋	OH^- (*aq*)	+	NH_4^+ (*aq*)
I	0.33 M		\		0		0
C	−x		\		+x		+x
E	0.33−x		\		x		x

$$1.8 \times 10^{-5} = x^2 / 0.33$$

$$x = 0.00244 \text{ M} = [OH^-]$$

$$\text{pOH} = -\log(0.00244) = 2.613$$

$pH = 14 - 2.613 =$ **11.39** (This pH makes sense as it is basic)

II. Only the **Weak Acid** remaining (Leads to SCENARIO 3)

For example, say a question asks about the pH of a 100 mL solution of 1 M NH_3 that is titrated with 200 mL of 0.5 M HCl ($K_a = 5.5 \times 10^{-10}$ of NH_4^+):

1. Find mols of each reactant to determine the limiting reagent

$$HCl\ (aq) \rightarrow H^+\ (aq) + Cl^-\ (aq)$$

0.5 M	0.5 M	(Negl.)

H^+ from HCl: 0.2 L x 0.5 mol/L = **0.1 mol** H^+

Weak bases do not completely dissociate, so we will calculate the mols only for NH_3: 0.1 L x 1.0 mol/L = **0.1 mol** NH_3

HCl and NH_3 are the limiting reactant (equal moles)

2. Write the reaction and fill in the ICF table

	$NH_3\ (aq)$	+ $H^+\ (aq)$	$\rightarrow NH_4^+\ (aq)$
I	0.1 mol	0.1 mol	0
C	−0.1 mol	−0.1 mol	+0.1 mol
F	0	0	0.1 mol

Only weak acid (NH_4^+) remains, so we will follow the steps in **SCENARIO 3** to determine the pH of this solution. But first we need the concentration with the NEW VOLUME because we will be using an **ICE table**

$[NH_4^+]$: 0.1 mol / (0.1 L + 0.2 L) = **0.33 M**

3. Follow the steps in **SCENARIO 3**

	$NH_4^+\ (aq)$	+ $H_2O\ (\ell)$	$\leftrightharpoons H_3O^+\ (aq)$	+ $NH_3\ (aq)$
I	0.33 M	\	0	0
C	−x	\	+x	+x
E	0.33−x	\	x	x

$$5.5 \times 10^{-10} = x^2 / 0.33$$

$$x = 0.0000135\ M = [H^+]$$

$pH = -\log(0.0000135) = 4.87$ (This pH makes sense as it is acidic)

III. Both **Weak Acid** & **Weak Base** remaining (Leads to SCENARIO 4)

For example, say a question asks about the pH of a 100 mL solution of 1 M NH_4^+ that is titrated with 50 mL of 0.5 M NaOH ($K_a = 5.5 \times 10^{-10}$ of NH_4^+):

1. Find mols of each reactant to determine the limiting reagent

$NaOH\ (aq)$	$\rightarrow$	$Na^{+}\ (aq)$	$+$	$OH^{-}\ (aq)$
0.5 M		(Negl.)		0.5 M

OH^{-} from NaOH: 0.05 L x 0.5 mol/L = **0.025 mol OH^{-}**

Weak acids do not completely dissociate, so we will calculate the mols only for NH_4^{+}: 0.1 L x 1.0 mol/L = **0.1 mol NH_4^{+}**

NaOH is the limiting reactant (less moles)

2. Write the reaction and fill in the ICF table

	$OH^{-}\ (aq)$	$+\ NH_4^{+}\ (aq)$	$\rightarrow\ NH_3\ (aq)$	$+\ H_2O\ (\ell)$
I	0.025 mol	0.1 mol	0	\
C	−0.025 mol	−0.025 mol	+0.025 mol	\
F	0	0.075 mol	0.025 mol	\

Both weak base (NH_3) and weak acid (NH_4^{+}) remain, so we will follow the steps in **SCENARIO 4** to determine the pH of this solution. As this would be a **buffer solution**. First, we need the concentrations of each.

$[NH_3]$: 0.025 mol / (0.1 L + 0.05 L) = **0.167 M**

$[NH_4^{+}]$: 0.075 mol / (0.1 L + 0.05 L) = **0.5 M**

3. Follow the steps in **SCENARIO 4**

$$pK_a = -\log(K_a) = -\log(5.5 \times 10^{-10}) = \mathbf{9.26}$$

$$pH = 9.26 + \log\left([0.167\ M]/[0.5\ M]\right) = 8.78$$

** If the concentration of the conjugate pair is equal, the **pH = p*Ka***. This is known as the *half-equivalence* point and occurs because:

$$[Base]/[Acid] = 1$$
$$pH = pK_a + \log(1) = pK_a + 0$$
Thus, **pH = p*Ka***

Practice below by finding the pH of a 150 mL solution of 1.25 M NH_3 that is titrated with 75 mL of 1 M HCl: