Basics of calculating pH

1. Find the pH of 0.07 M HCl.

2. Find the pH of 0.2 M propanoic acid ($K_a = 10^{-4.87}$)

3. Find the pH of 0.4 M $(CH_3)_3N$ ($K_b = 10^{-4.20}$)

4. Find the pH of 0.3 M $CH_3COO^-Na^+$ (K_a of ethanoic acid is $10^{-4.76}$)

5. Find the pH of 0.7 M NH_4NO_3 (K_b of ammonia is $10^{-4.75}$)

Relative Strengths of Acids and Bases

6. Which is the strongest acid here?

Name	Formula	pK_a
methanoic	$HCOOH$	3.75
ethanoic	CH_3COOH	4.76
propanoic	CH_3CH_2COOH	4.87
2,2-dimethylpropanoic	$(CH_3)_3CCOOH$	5.03

7. Which is the strongest base here?

Name	Formula	pK_b
ammonia	NH_3	4.75
methylamine	CH_3NH_2	3.34
trimethylamine	$(CH_3)_3N$	4.20
phenylamine	$C_6H_5NH_2$	9.13

Conjugate Acids and Bases

8. Write the chemical formula (with proper charge) for each of the following:

a) conjugate acid of HSO_4^-

b) conjugate base of HSO_4^-

c) conjugate base of $CH_3(CH_2)_6COOH$

d) conjugate acid of $(C_6H_5)_3N$

e) conjugate acid of O^{2-}

f) conjugate base of $HC{\equiv}CH$

g) conjugate base of H_2NCH_2COOH

h) conjugate acid of H_2NCH_2COOH

Buffers

9. Does each solution constitute a *buffer*? Explain your answer.

a) One mole of HCl mixed with two moles of NaOH

b) One mole of HCl mixed with half a mole of NaOH

c) One mole of HCl mixed with two moles of NH_3

d) One mole of HCl mixed with one mole of NH_3

e) 0.1 M NH_3 with 0.1 M NH_4Cl

f) 0.1 M $HC_2H_3O_2$ with 0.1 M $NaC_2H_3O_2$

g) 0.1 M HNO_3 with 0.1 M $NaNO_3$

h) 10 mL of 0.1 M H_2SO_3 mixed with 10 mL of 0.05 M KOH

i) 0.1 M H_2SO_4 with 0.1 M $NaHSO_4$

j) 0.1 M $NaHSO_4$ with 0.1 M Na_2SO_4

10. What is the pH of each buffer solution here?

a) 100 mL solution, containing 0.2 M $HF_{(aq)}$ and 0.15 M $NaF_{(aq)}$

K_a of HF is 6.6×10^{-4}

b) 100 mL solution, made from 50 mL of 0.3 M $HClO_{(aq)}$ and 50 mL of 0.1 M $KOH_{(aq)}$.

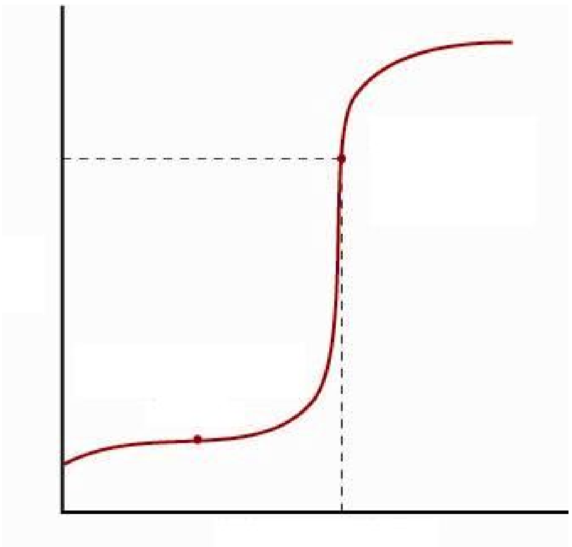
K_a of HClO is 2.9×10^{-8}

c) A solution that contains 0.5 mol of methylamine, which has 0.15 mol of HBr added to it.

K_b for methylamine is 4.4×10^{-4}

Titration Curves

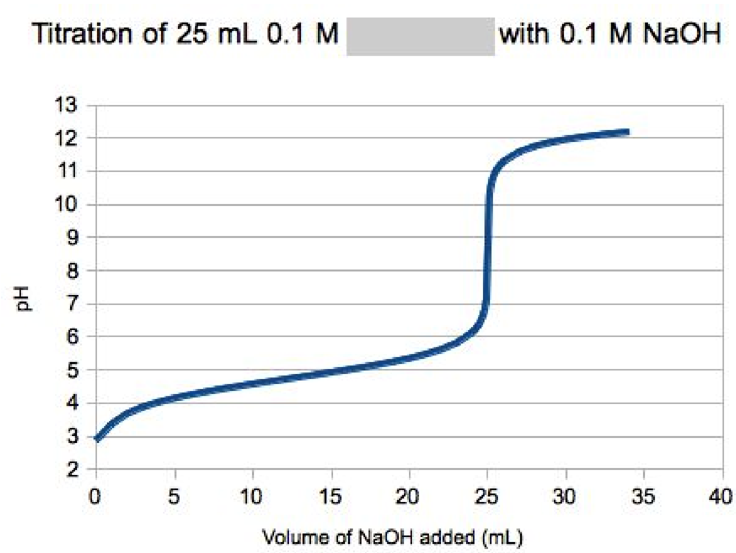
11. Label this titration curve. It needs:
    - Labels on the *x* and *y* axes
    - Labels for each of the dots
    - A title: Whether each solution (acid and base) are weak/strong, and what's the "titrant"?



12.

a) Identify the pKa of the acid that was titrated here.

b) What would be a good indicator for this titration? Explain your choice.



Titration Calculations

13. In the titration of 25.00 mL of 0.100 M HCl with 0.100 M NaOH, what is the pH of the solution after 15.00 mL of the standard NaOH solution has been added?

14. When titrating 20.00 mL of 0.12 M HCN with 0.08 M NaOH, you have reached the equivalence point when you add 30.00 mL of the NaOH. What is the pH at this point?
K_a for HCN is 6.2×10^{-10}.

15. 1.0 g of an unknown acid was dissolved in 100 mL of water. 25.00 mL of this solution was titrated with 0.100 M KOH, and the experimental data is shown below.
What is the identity of the acid?

Trial	**Burette Readings (mL)**	
	Initial	**Final**
1	0.02	20.58
2	12.11	32.69
3	14.32	34.66

Answers

1. Find the pH of 0.07 M HCl.

For a strong acid, $[H_3O^+]$ = concentration of acid
pH = $-\log[H_3O^+] = -\log(0.07) =$ **1.15**

2. Find the pH of 0.2 M propanoic acid ($K_a = 10^{-4.87}$)

For a weak acid, you may want to create an ICE table.
$K_a = \frac{x^2}{0.2-x} \approx \frac{x^2}{0.2}$ Since C/K = $0.2 / 10^{-4.87} = 14826 >> 1000$
$x^2 = (0.2)(10^{-4.87}) = 2.697 \times 10^{-6}$
$x = 1.64 \times 10^{-3}$

So pH = $-\log(1.64 \times 10^{-3}) =$ **2.78**

3. Find the pH of 0.4 M $(CH_3)_3N$ ($K_b = 10^{-4.20}$)

For a weak base, you may want to create an ICE table.
$K_b = \frac{x^2}{0.4-x} \approx \frac{x^2}{0.4}$ Since C/K = $0.4 / 10^{-4.20} = 6339 >> 1000$
$x^2 = (0.4)(10^{-4.20}) = 2.524 \times 10^{-5}$
$x = 5.02 \times 10^{-3}$

So pOH = $-\log(5.02 \times 10^{-3}) = 2.30$
pH = 14 – 2.30 = **11.70**

4. Find the pH of 0.3 M $CH_3COO^-Na^+$ (K_a of ethanoic acid is $10^{-4.76}$)

We need the K_b of ethanoate (the conjugate base). $K_b = 10^{-14} / K_a = 10^{-14} / 10^{-4.76} = 5.75 \times 10^{-10}$.

$K_b = \frac{x^2}{0.3-x} \approx \frac{x^2}{0.3}$ Since C/K = $0.3 / (5.75 \times 10^{-10}) >> 1000$
$x^2 = (0.3)(5.75 \times 10^{-10}) = 1.726 \times 10^{-10}$
$x = 1.31 \times 10^{-5}$

So pOH = $-\log(1.31 \times 10^{-5}) = 4.88$
pH = 14 – 4.88 = **9.12**

5. Find the pH of 0.7 M NH_4NO_3 (K_b of ammonia is $10^{-4.75}$)

We need the K_a of NH_4^+ (the conjugate acid of NH_3) and so $K_a = 10^{-14} / 10^{-4.75} = 5.62 \times 10^{-10}$.

$K_a = \frac{x^2}{0.7-x} \approx \frac{x^2}{0.7}$ Since C/K = $0.7 / (5.62 \times 10^{-10}) >> 1000$
$x^2 = (0.7)(5.62 \times 10^{-10}) = 3.936 \times 10^{-10}$
$x = 1.98 \times 10^{-5}$

So pH = $-\log(1.98 \times 10^{-5}) =$ **4.70**

6. Strongest acid → Find the lowest pK_a or highest K_a

Name	Formula	pK_a	
methanoic	$HCOOH$	3.75	Strongest acid of these four
ethanoic	CH_3COOH	4.76	
propanoic	CH_3CH_2COOH	4.87	
2,2-dimethylpropanoic	$(CH_3)_3CCOOH$	5.03	Weakest acid of these four

7. Looking for the strongest base of a list? → Find the lowest pK_b, or highest K_b

Name	Formula	pK_b	
ammonia	NH_3	4.75	
methylamine	CH_3NH_2	3.34	Strongest base of these four
trimethylamine	$(CH_3)_3N$	4.20	
phenylamine	$C_6H_5NH_2$	9.13	Weakest base of these four

8.a) conjugate acid of HSO_4^- H_2SO_4

b) conjugate base of HSO_4^- SO_4^{2-}

c) conjugate base of $CH_3(CH_2)_6COOH$ $CH_3(CH_2)_6COO^-$

d) conjugate acid of $(C_6H_5)_3N$ $(C_6H_5)_3NH^+$

e) conjugate acid of O^{2-} OH^-

f) conjugate base of $HC{\equiv}CH$ $HC{\equiv}C^-$

g) conjugate base of H_2NCH_2COOH $H_2NCH_2COO^-$

h) conjugate acid of H_2NCH_2COOH $^+H_3NCH_2COOH$

9. Does each solution constitute a *buffer*? Explain your answer.

a) One mole of HCl mixed with two moles of NaOH No. Strong base remains.

b) One mole of HCl mixed with half a mole of NaOH No. Strong acid remains.

c) One mole of HCl mixed with two moles of NH_3 Yes. Mixture is half NH_3, half NH_4^+

d) One mole of HCl mixed with one mole of NH_3 No. Mixture is all NH_4^+

e) 0.1 M NH_3 with 0.1 M NH_4Cl Yes. half NH_3 and half NH_4^+

f) 0.1 M $HC_2H_3O_2$ with 0.1 M $NaC_2H_3O_2$ Yes. Half weak acid, half conj base

g) 0.1 M HNO_3 with 0.1 M $NaNO_3$ — No. HNO_3 is strong → Not a buffer

h) 10 mL of 0.1 M H_2SO_3 mixed with 10 mL of 0.05 M KOH — Yes. Half H_2SO_3 and half conj base

i) 0.1 M H_2SO_4 with 0.1 M $NaHSO_4$ — No. H_2SO_4 is strong.

j) 0.1 M $NaHSO_4$ with 0.1 M Na_2SO_4 — Yes. Half HSO_4^- and half conj base

10. What is the pH of each buffer solution here?

a) 100 mL solution, containing 0.2 M $HF_{(aq)}$ and 0.15 M $NaF_{(aq)}$

$$pH = pKa + log(\frac{[base]}{[acid]})$$
$$pH = -log(6.6\times10^{-4}) + log(\frac{0.15}{0.2}) = 3.18 + -0.12 = 3.06$$

b) 100 mL solution, made from 50 mL of 0.3 M $HClO_{(aq)}$ and 50 mL of 0.1 M $KOH_{(aq)}$.

$$pH = pKa + log(\frac{[base]}{[acid]})$$

$$pH = -log(2.9\times10^{-8}) + log(\frac{\frac{0.1}{2}}{\frac{0.3}{2}}) = 7.54 + -0.48 = 7.06$$

c) A solution that contains 0.5 mol of methylamine, which has 0.15 mol of HBr added to it.

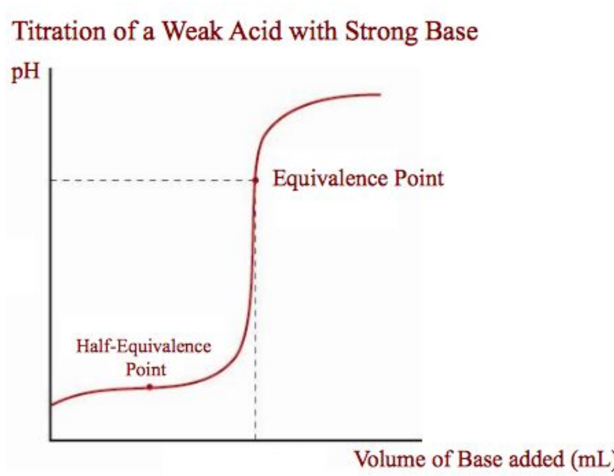
K_a of methylamine-H^+ is $K_w/K_b = 10^{-14} / K_b = 10^{-14} / (4.4 \times 10^{-4}) = 2.27 \times 10^{-11}$

$$pH = pKa + log(\frac{[base]}{[acid]})$$

$$pH = -log(2.27\times10^{-11}) + log(\frac{0.5-0.15}{0+0.15})$$
$= 10.64 + 0.37$
= 11.01

Titration Curves

11. Label this titration curve. It needs:
Labels on the *x* and *y* axes
Labels for each of the dots
A title: Whether each solution (acid and base) are weak/strong, and what's the "titrant"?



12.
a) Identify the acid that was titrated here.
b) What would be a good indicator for this titration? Explain your choice.

a) The pH at halfway-to-equivalence is about 4.75
Therefore, the pKa of the acid is 4.75
Therefore, the Ka of the acid is $10^{-4.75} = 1.8 \times 10^{-5}$ ... probably acetic acid.

b) Anything that changes pH in the slightly basic range (from 8 to 10).
Examples include: Thymol blue (Yellow to blue from 8.0 - 9.6)
Phenolphthalein (Colorless to pink from 8.2 - 10.0)

13. In the titration of 25.00 mL of 0.100 M HCl with 0.100 M NaOH, what is the pH of the solution after 15.00 mL of the standard NaOH solution has been added?

$n_{HCl} = CV = (0.100\text{ M})(0.025\text{ L}) = 0.0025\text{ mol}$
$n_{NaOH} = CV = (0.100\text{ M})(0.015\text{ L}) = 0.0015\text{ mol}$

Therefore, NaOH is limiting and HCl is in excess. There will be 0.0010 mol of HCl remaining in the 0.04 L mixture. So [HCl] = 0.0010 mol / 0.04 L = 0.025 M

$pH = -\log[H_3O^+] = -\log(0.025) = 1.60$

14. When titrating 20.00 mL of 0.12 M HCN with 0.08 M NaOH, you have reached the equivalence point when you add 30.00 mL of the NaOH. What is the pH at this point?
K_a for HCN is 6.2×10^{-10}.

Just checking to make sure it's the equivalence point:
$N_{HCN} = (0.020\text{ L})(0.12\text{ M}) = 0.0024\text{ mol}$
$n_{NaOH} = (0.030\text{ L})(0.08\text{ M}) = 0.0024\text{ mol}$

What's left in the solution at the equivalence point?

$HCN + NaOH \rightarrow H_2O + Na^+ + CN^-$

None of the HCN or NaOH remains. H_2O and Na^+ don't affect pH. Therefore, finding the pH of the solution at equivalence is like asking for the pH of 0.0024 mol CN^- in 50 mL of water.

K_b for CN^- is $10^{-14}/(6.2 \times 10^{-10}) = 1.613 \times 10^{-5}$.
$[CN^-]$ is 0.0024 mol / 0.050 L = 0.048 M

$K_b = \frac{x^2}{0.048-x} \approx \frac{x^2}{0.048}$ Since $C/K = 0.048 / (1.613 \times 10^{-5}) >> 1000$
$x^2 = (0.048)(1.613 \times 10^{-5}) = 7.742 \times 10^{-7}$
$x = 8.80 \times 10^{-4}$

So $pOH = -\log(8.80 \times 10^{-4}) = 3.06$ and $pH = 14 - 3.06 =$ **10.94**

15. 1.0 g of an unknown acid was dissolved in 100 mL of water. 25.00 mL of this solution was titrated with 0.100 M KOH, and the experimental data is shown below.
What is the identity of the acid?

Trial	Burette Readings (mL)	
	Initial	Final
1	0.02	20.58
2	12.11	32.69
3	14.32	34.66

Volumes Needed Trial 1: 20.56 mL
Trial 2: 20.58 mL
Trial 3: 20.34 mL
Average: 20.49 mL of 0.100 M KOH

Amount of KOH added: n = CV = (0.100 M)(0.02049 L) = 0.002049 mol
Amount of acid in the 25.00 mL sample: 0.002049 mol (definition of equivalence)
Amount of acid in the 1.0 g sample: 0.002049 mol × 4 = 0.008196 mol

Molar mass of acid: M = m/n = (1.0 g) / (0.008196 mol) = 122 g/mol
This corresponds to **benzoic acid**, C_6H_5COOH (aka $C_7H_6O_2$)