

AP Practice Problems
Acid/Base Equilibria

1984 C

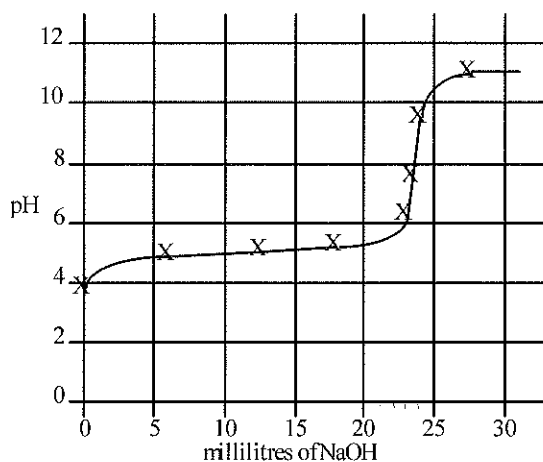
Discuss the roles of indicators in the titration of acids and bases. Explain the basis of their operation and the factors to be considered in selecting an appropriate indicator for a particular titration.

1986 A

In water, hydrazoic acid, HN_3 , is a weak acid that has an equilibrium constant, K_a , equal to 2.8×10^{-5} at 25°C . A 0.300 litre sample of a 0.050 molar solution of the acid is prepared.

- Write the expression for the equilibrium constant, K_a , for hydrazoic acid.
- Calculate the pH of this solution at 25°C .
- To 0.150 litre of this solution, 0.80 gram of sodium azide, NaN_3 , is added. The salt dissolved completely. Calculate the pH of the resulting solution at 25°C if the volume of the solution remains unchanged.
- To the remaining 0.150 litre of the original solution, 0.075 litre of 0.100 molar NaOH solution is added. Calculate the $[\text{OH}^-]$ for the resulting solution at 25°C .

1988 D



A 30.00 millilitre sample of a weak monoprotic acid was titrated with a standardized solution of NaOH. A pH meter was used to measure the pH after each increment of NaOH was added, and the curve above was constructed.

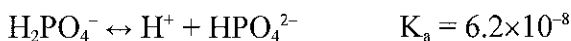
- Explain how this curve could be used to determine the molarity of the acid.
- Explain how this curve could be used to determine the dissociation constant K_a of the weak monoprotic acid.
- If you were to repeat the titration using an indicator in the acid to signal the endpoint, which of the following indicators should you select? Give the reason for your choice.

| | |
|-----------------|---------------------------|
| Methyl red | $K_a = 1 \times 10^{-5}$ |
| Cresol red | $K_a = 1 \times 10^{-8}$ |
| Alizarin yellow | $K_a = 1 \times 10^{-11}$ |
- Sketch the titration curve that would result if the weak monoprotic acid were replaced by a strong monoprotic acid, such as HCl of the same molarity. Identify differences between this titration curve and the curve shown above.

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4th Question

The equations and constants for the dissociation of three different acids are given below.



- From the systems above, identify the conjugate pair that is best for preparing a buffer with a pH of 7.2. Explain your choice.
- Explain briefly how you would prepare the buffer solution described in (a) with the conjugate pair you have chosen.
- If the concentrations of both the acid and the conjugate base you have chosen were doubled, how would the pH be affected? Explain how the capacity of the buffer is affected by this change in concentrations of acid and base.
- Explain briefly how you could prepare the buffer solution in (a) if you had available the solid salt of the only one member of the conjugate pair and solution of a strong acid and a strong base.

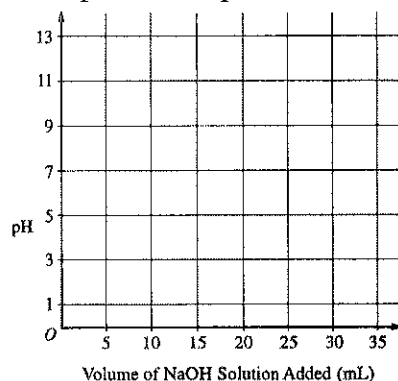
1998 D (Required)

[repeated in lab procedures section]

An approximately 0.1-molar solution of NaOH is to be standardized by titration. Assume that the following materials are available.

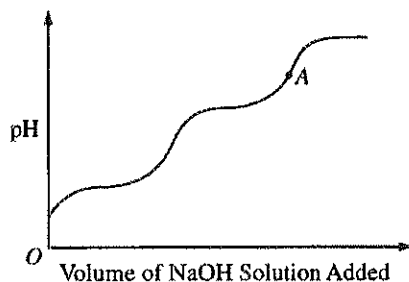
- Clean, dry 50 mL buret
- 250 mL Erlenmeyer flask
- Wash bottle filled with distilled water
- Analytical balance
- Phenolphthalein indicator solution
- Potassium hydrogen phthalate, KHP, a pure solid monoprotic acid (to be used as the primary standard)

- Briefly describe the steps you would take, using the materials listed above, to standardize the NaOH solution.
- Describe (*i.e.*, set up) the calculations necessary to determine the concentration of the NaOH solution.
- After the NaOH solution has been standardized, it is used to titrate a weak monoprotic acid, HX. The equivalence point is reached when 25.0 mL of NaOH solution has been added. In the space provided at the right, sketch the titration curve, showing the pH changes that occur as the volume of NaOH solution added increases from 0 to 35.0 mL. Clearly label the equivalence point on the curve.



AP Practice Problems
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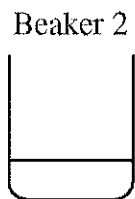
- (d) Describe how the value of the acid-dissociation constant, K_a , for the weak acid HX could be determined from the titration curve in part (c).
- (e) The graph below shows the results obtained by titrating a different weak acid, H_2Y , with the standardized NaOH solution. Identify the negative ion that is present in the highest concentration at the point in the titration represented by the letter A on the curve.



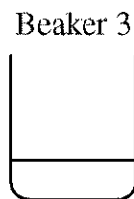
1. Each of three beakers contains 25.0 mL of a 0.100 M solution of HCl, NH_3 , or NH_4Cl , as shown above. Each solution is at 25° C.



0.100 M HCl



0.100 M NH_3



0.100 M NH_4Cl

2. (a) Determine the pH of the solution in beaker 1. Justify your answer.
- (b) In beaker 2, the reaction $NH_3(aq) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq)$ occurs. The value of K_b for $NH_3(aq)$ is 1.8×10^{-5} at 25° C.
- Write the K_b expression for the reaction of $NH_3(aq)$ with $H_2O(l)$.
 - Calculate the $[OH^-]$ in the solution in beaker 2.
- (c) In beaker 3, the reaction $NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$ occurs.
- Calculate the value of K_a for $NH_4^+(aq)$ at 25° C.
 - The contents of beaker 2 are poured into beaker 3 and the resulting solution is stirred. Assume that volumes are additive. Calculate the pH of the resulting solution.
- (d) The contents of beaker 1 are poured into the solution made in part (c)(ii). The resulting solution is stirred. Assume that volumes are additive.
- Is the resulting solution an effective buffer? Justify your answer.
 - Calculate the final $[NH_4^+]$ in the resulting solution at 25° C.

Last Question

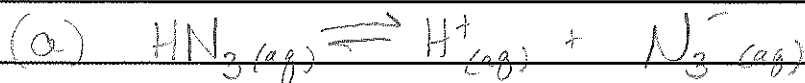
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ADDITIONAL PAGE FOR ANSWERING QUESTION 5

1984 C

Indicators are used to show when the equivalence point of a titration is reached. Because there is a sharp increase in pH as the equivalence point is reached the color of the indicator will change. The indicator is one color when it is protonated (HIn) and a different color when it is non-protonated (In^-). To choose an appropriate indicator, it is necessary for the pK_a of the indicator to be similar to the pH at the equivalence point.

1986 A



$$K_a = \frac{[\text{H}^+][\text{N}_3^-]}{[\text{HN}_3]}$$



$$I \quad 0.050 \text{ M} \quad 0 \quad 0$$

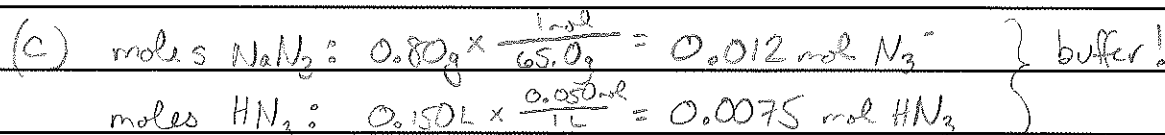
$$C \quad -x \quad +x \quad +x$$

$$E \quad 0.050 - x \quad x \quad x$$

$$K_a = 2.8 \times 10^{-5} = \frac{x^2}{0.050 - x} \approx \frac{x^2}{0.050 \text{ M}}$$

$$x = 1.18 \times 10^{-3} \text{ M} = [\text{H}^+]$$

$$\text{pH} = 2.93$$



$$\text{pH} = \text{pK}_a + \log \frac{[\text{N}_3^-]}{[\text{HN}_3]} = 4.55 + \log \frac{0.012}{0.0075} = 4.75$$

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ADDITIONAL PAGE FOR ANSWERING QUESTION 5

1986A (cont)



I 0.0075 mol 0.0075 mol

0

$[\text{N}_3^-] = \frac{0.0075 \text{ mol}}{0.225 \text{ L}}$

C -0.0075 -0.0075

+0.0075

$[\text{N}_3^-] = 0.0333 \text{ M}$

F 0 mol 0 mol

0.0075 mol



$K_b = \frac{[\text{HN}_3][\text{OH}^-]}{[\text{N}_3^-]} = \frac{K_w}{K_a} = 3.57 \times 10^{-10}$

I 0.0333 M

0

0

C -x

+x

+x

$3.57 \times 10^{-10} = \frac{x^2}{0.0333}$

E 0.0333 - x

x

x

$x = 3.45 \times 10^{-6} \text{ M OH}^-$

$\text{pOH} = 5.46$

$\text{pH} = 8.54$

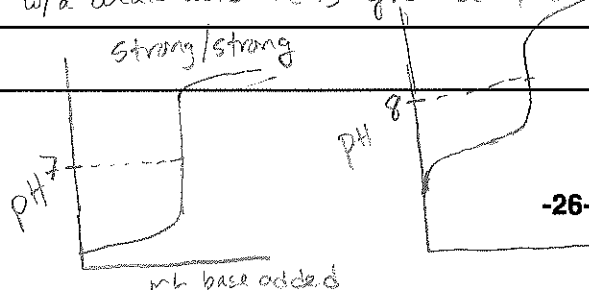
1988D

(a) The curve can be used to determine the equivalence point by looking for the most vertical increase in pH. At the equivalence point $M_A V_A = M_B V_B$. In this titration, you would know V_A , M_B , and V_B , so you could solve for M_A .

(b) The pH at the half-way point ($\sim 11.75 \text{ mL NaOH}$) would be equal to the pK_a of the weak acid. The $K_a = 10^{-\text{pK}_a}$.

(c) Cresol Red should be used. The K_a of the indicator should be similar to the pH at the endpoint of the titration. This ensures that the color change occurs at a pH similar to the pH at the equivalence point.

(d) For a strong acid titration, the pH at the eq. point will be 7 whereas w/a weak acid it is greater than 7. The strong also has a more constant



slope at the beginning of the curve whereas the weak has a sharper increase at the beginning, but then it levels out more in the buffering region.

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ADDITIONAL PAGE FOR ANSWERING QUESTION 5

4th Question

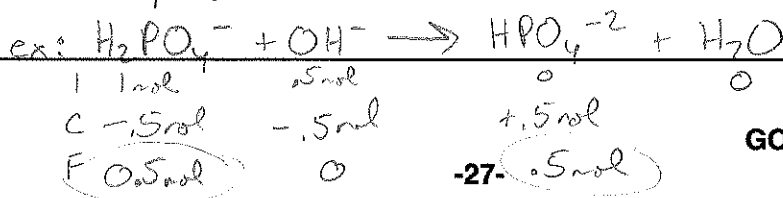
(a) $\text{H}_2\text{PO}_4^- / \text{HPO}_4^{2-}$ is the best conjugate pair because the pH of the buffer should be similar to the pKa of the acid in the conjugate pair.

$$\text{pKa} = -\log(6.2 \times 10^{-8}) = 7.2$$

(b) Because the $\text{pH} = \text{pKa}$ in this case, I would dissolve an equal number of moles of a H_2PO_4^- salt (such as NaH_2PO_4) and a HPO_4^{2-} salt (Na_2HPO_4 ← such as) in a sample of distilled water.

(c) The pH would not change if the concentrations were doubled. pH of a buffer will change if the RATIO of the concentrations change, but not just by doubling both since the ratio stays the same in that case. The buffer with the doubled concentrations will have a greater buffering capacity though because it will be able to neutralize twice as much added strong acid or base.

(d) If the salt NaH_2PO_4 was available, I could dissolve a certain # of moles of that salt in water. I would then add half that number of moles of the strong base to the solution. This would result in a solution with an equal concentration of H_2PO_4^- and HPO_4^{2-} .



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1998 D

- (a)
- 1) Measure out 0.500g of KHP on the analytical balance
 - 2) Dissolve the KHP in 20.0ml H₂O in the erlenmeyer flask
 - 3) Add 3 drops of phenolphthalein to the flask
 - 4) Fill the buret with the NaOH solution. (Make sure the tip is filled)
 - 5) Record the volume of NaOH in the buret.
 - 6) Titrate the KHP solution w/ the NaOH solution until a light pink color appears + persists for ^{at least} 15 seconds
 - 7) Record the volume of NaOH in the buret

(b)

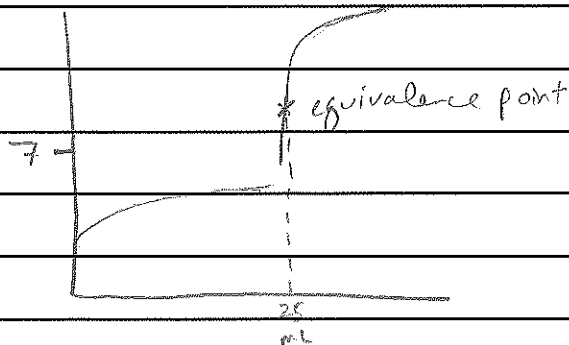
- 1) Convert 0.500g of KHP to moles of KHP by dividing by the molar mass of KHP.

$$0.500g \text{ KHP} \times \frac{1 \text{ mol KHP}}{\text{Molar Mass KHP}} = \text{mol KHP}$$

- 2) mol KHP = mol NaOH at equivalence point
- 3) Divide mol NaOH by the volume of NaOH solution needed to reach the equivalence point.

$$M \text{ NaOH} = \frac{\text{mol NaOH}}{\text{Volume NaOH}}$$

(c)



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ADDITIONAL PAGE FOR ANSWERING QUESTION 6

(1998D cont)

(d) The $pK_a = pH$ at the half-way point.

$$10^{-pK_a} = K_a \text{ for the acid}$$

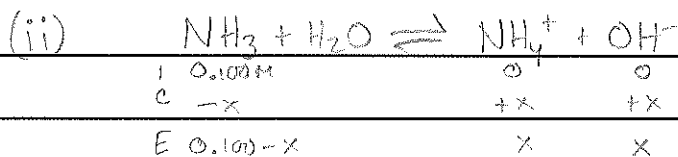
(e) y^{-2}

(a) $pH = -\log[H^+]$ b/c this is a strong acid $[H^+] = 0.10M$
 $pH = 1.000$

LAST QUESTION

(b)

$$(i) K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$



$$K_b = 1.8 \times 10^{-5} = \frac{x^2}{0.100M}$$

$$x = 1.3 \times 10^{-3} M OH^-$$

(c) (i) $K_a \text{ for } NH_4^+ = \frac{K_w}{K_b \text{ for } NH_3} = 5.6 \times 10^{-10}$

(ii) Results in a buffer where $[NH_3] = [NH_4^+]$

Therefore $pH = pK_a = -\log(5.6 \times 10^{-10}) = 9.26$

(d) (i) The resulting solution is not an effective buffer.

The addition of the HCl reacts w/ the NH_3 to form NH_4^+ .

This results in a solution that is primarily NH_4^+ and no longer works as a buffer.

