

Hydrolysis

A salt of a weak acid acts as a weak base in water solution. In other words the anion of a weak acid will remove H_3O^+ ions from solution to form the unionized acid, which, being weak, does not ionize again to a very great extent. Then, of course, the solution will be basic, since an excess of OH^- ions is left. Similarly, a salt of a weak base will form an acidic solution.

A typical hydrolysis problem asks for the pH, pOH, $[\text{H}^+]$ or $[\text{OH}^-]$ of a solution of a salt. The solution of such a problem can be carried out like any other acid-base equilibrium, using the appropriate K_a and/or K_b for the ion(s) present.

Procedure:

- (1) Examine the salt that is present:
 - (a) is the anion the conjugate base of a weak acid?
 - (b) is the cation the conjugate acid of a weak base?(See, for example, table 13-2, Goldberg and Dillard)
- (2) Write an equation showing the formation of the weak acid or weak base from the reaction of the ion with water, for each ion found in (1) to be the conjugate of a weak species.
- (3) If an equilibrium constant is not given or requested for the reaction(s) in (2), look it up. Be sure you know which K you have, K_a or K_b , and what equilibrium it applies to.
- (4) If the available K is not the right one for the equation in (2), use the relationship, $K_w = K_a \times K_b$, which refers to the constants for conjugate acids and bases, to convert to the correct K.
- (5) Now proceed as you would for any equilibrium problem. (See, for example, the 'Equilibrium Constant Calculations' Approach Sheet.)

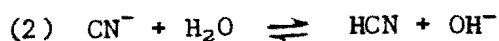
Examples which are worked out on the following pages are:

- (A) Find the pH of a 1.0 M. solution of NaCN.
- (B) What is the pH of a 0.1 M. solution of $\text{C}_5\text{H}_5\text{NHCl}$?
- (C) What is the pH of a solution 0.1 M. in KCl?

We recommend that you try to work these problems before you look at the solutions given on the following pages.

Solutions:

- (A) (1) From the table we find that HCN, the conjugate acid of CN^- is weak, with $K_a = 7.2 \times 10^{-10}$.
We know NaOH is strong, since Na is an alkali metal.



- (3) The equation in (2) is for CN^- acting as a base (taking H^+ from water) so it has a K_b . Since CN^- is the conjugate base of HCN, we know $K_b = K_w/K_a$ or:

$$K_b = \frac{10^{-14}}{7.2 \times 10^{-10}} = 1.39 \times 10^{-5} \quad \text{OR } 1.4 \times 10^{-5}$$

1.0 M is the initial concentration of CN^-

Let $X = [\text{OH}^-]$ at equilibrium.

There is no HCN present before reaction, so none can be consumed.

Assume that OH^- is zero initially and none will be consumed.

No additional CN^- can be produced, since nothing is present which can react to form it.

	CN^-	+	H_2O	\rightleftharpoons	HCN	+	OH^-
I	1.0 M		--		0		0
C	X		--		0		0
P	0		--		X		X
E	1.0 - X		--		X		X

Summing the OH^- column: $P = X$

Hence stoichiometry says: X of $[\text{HCN}]$ produced also and X of $[\text{CN}^-]$ must be consumed.

Now sum CN^- : $E = 1.0 - X$

and HCN: $E = X$

$$K_b = 1.4 \times 10^{-5} = \frac{[\text{HCN}][\text{OH}^-]}{[\text{CN}^-]} = \frac{X^2}{1.0 - X}$$

Assume $1.0 - X = 1.0$ and CHECK LATER.

$$X^2 = 1.4 \times 10^{-5} = 14 \times 10^{-6}, \text{ so } X = 3.7 \times 10^{-3}$$

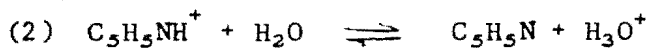
Check approximation: $1.0 - .0037 = 1.0$ to our significant figures.

Also note $1.0 \times 10^{-7} + 3.7 \times 10^{-3} = 3.7 \times 10^{-3}$ so the initial $[\text{OH}^-]$ was not a problem.

Now calculate pH: $[\text{OH}^-] = 3.7 \times 10^{-3}$ so $\text{pOH} = -\log 3.7 + 3 = -.57 + 3 = 2.43$
 $\text{pOH} + \text{pH} = 14$, so $\text{pH} = 14 - 2.43 = 11.57$, or 11.6

(B) From the table we find that C_5H_5NH is a weak base, with $K_b = 2.3 \times 10^{-9}$.

(1) The ions involved are $C_5H_5NH^+$ and Cl^- .



(3) For this equation we need K_a , but we have K_b for the conjugate so:

$$K_a = \frac{1.0 \times 10^{-14}}{2.3 \times 10^{-9}} = 4.35 \times 10^{-6}$$

0.1 M is the initial concentration of $C_5H_5NH^+$

There is no C_5H_5N initially present, so none consumed.

Assume the H_3O^+ initially present is small compared with the amount produced by the hydrolysis. (check this when it is calculated)

No H_3O^+ consumed, no $C_5H_5NH^+$ produced.

Let $X = [H_3O^+]$ present at equilibrium.

	$C_5H_5NH^+$	+	H_2O		C_5H_5N	+	H_3O^+
I	0.1		--		0		0
C	X		--		0		0
P	0		--		X		X
E	0.1 - X		--		X		X

Now sum the H_3O^+ column: $P = X$

Since X of $[H_3O^+]$ is produced, X of C_5H_5N is also, and X of $C_5H_5NH^+$ must be consumed.

Now sum the other columns:

$$C_5H_5NH^+: E = 0.1 - X$$

$$C_5H_5N: E = X$$

$$K_b = 4.35 \times 10^{-6} = \frac{[H_3O^+][C_5H_5N]}{[C_5H_5NH^+]} = \frac{X^2}{0.1 - X}$$

Assume $0.1 - X \approx 0.1$: $X^2 = 4.35 \times 10^{-7} = 43.5 \times 10^{-8}$, $X = 6.6 \times 10^{-4}$

Check assumptions: $1.0 \times 10^{-7} + 6.6 \times 10^{-4} = 6.6 \times 10^{-4}$ $0.1 - 6.6 \times 10^{-4} = .1$

$$pH = -\log 6.6 + 4 = -.82 + 4 = 3.2$$

(C) The ions involved are K^+ and Cl^- . The conjugate of K^+ would be KOH, but it is strong. The conjugate of Cl^- would be HCl, but it also is strong. Consequently, neither ion has a tendency to remove H^+ or OH^- from solution, and the pH will be unchanged, that is 1.0×10^{-7} , as in water.

Acid and Base Equilibria

problems involving acid-base equilibria usually involve finding the $[H_3O^+]$ or $[OH^-]$ concentrations, but might also or instead ask for the concentration of some other ion present, which would be affected by the acid-base equilibrium.

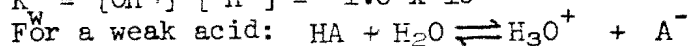
Relationships:

$$pH = -\log [H_3O^+] \quad \text{by definition of pH}$$

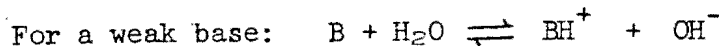
$$pOH = -\log [OH^-] \quad \text{by definition of pOH}$$

$$pK = -\log K \quad \text{for any } K.$$

$$K_w = [OH^-] [H^+] = 1.0 \times 10^{-14}$$



$$\text{Then: } K_a = \frac{[H_3O^+] [A^-]}{[HA]}$$



$$K_b = \frac{[BH^+] [OH^-]}{[B]}$$

Using the above relationships one can solve acid-base equilibrium problems using the method given for solving any equilibrium constant problem.

Method:

- (1) Determine what equilibrium is involved.
 - (a) for a strong acid or base only the self-ionization of water is involved.
 - (b) for a weak acid or base the partial ionization equilibrium is what you need. You may treat the species as 'weak' whenever you are given or asked for an equilibrium constant. (As in Table 13-2, Goldberg and Dillard, for example)
- (2) Write the chemical equation for that equilibrium.
- (3) Find a value for the equilibrium constant. (It may be possible to use either a K_a or a K_b , but be sure you don't use the wrong value for the equilibrium constant expression you choose.)
- (4) See the 'Equilibrium Constant Calculations' Approach Sheet for a method of proceeding from here.

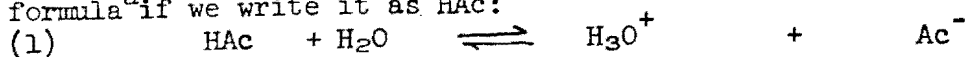
Examples:

- (A) What are $[H_3O^+]$ and pH for a 0.06 M. solution of acetic acid? $K_a = 1.8 \times 10^{-5}$.
- (B) What is the OH^- concentration of a 0.01 M. HCl solution?
- (C) Calculate K_a for HCN if a 1.0 M. solution has a pH of 4.57.
- (D) Find pOH of a 0.20 M. solution of CH_3NH_2 , where $K_b = 2.3 \times 10^{-10}$.

These examples are worked out on the following pages using the method given on the 'Equilibrium Constant Calculations' Approach Sheet. Try them yourself.

Solutions:

(A) Since K_a is given, acetic acid is weak. We don't need to know its chemical formula if we write it as HAc:



(2) I	0.06 M	----	0	0
C	X M		0	0
P	0		X M	X M
E	(0.06 - X)M		X M	X M

(3) 0.06 M is the initial (before ionization) concentration of HAc.

Let $X = [\text{H}^+]$ at equilibrium (as requested).

There is no $[\text{H}^+]$ or $[\text{OH}^-]$ present initially (we could use 10^{-7} for each from the water, but in an acid-base problem this is negligible compared to the acid or base contribution -- to check this we can compare the value we get for X when we finish.)

Because no H^+ or Ac^- is initially present, none can be consumed and no HAc can be produced, so put zeroes in those positions.

Note that we have ignored the concentrations of water, since we are using a water solution and therefore the amount of water present will not be affected appreciably by the reaction and is included in the K_a .

(4) The values will be inserted into the table in (1), but be sure you understand how each of them was arrived at: DON'T JUST STICK IN X'S ARBITRARILY!

(5) Complete the $[\text{H}^+]$: $X = 0 + P - 0 = P$

(6) Since X of H^+ is produced per liter, X of Ac^- must also be and X of HAc must be consumed.

(7) For $[\text{HAc}]$: $E = 0.06 + 0 - X = 0.06 - X$

For $[\text{Ac}^-]$: $E = 0 + X - 0 = X$

(8) $K_a = 1.8 \times 10^{-5} = \frac{[\text{H}^+][\text{Ac}^-]}{[\text{HAc}]}$

(9) $1.8 \times 10^{-5} = \frac{X^2}{0.06 - X}$

(10) Try assuming $X \ll 0.06$ to simplify this. Then:

$$X^2 = 0.06 \times 1.8 \times 10^{-5} = 1.08 \times 10^{-6}$$

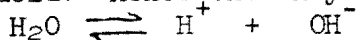
$$X = 1.04 \times 10^{-3} \text{ M or } 1.0 \times 10^{-3}$$

Check approximation: $0.06 - X = 0.06 - 0.001 = 0.059 \sim 0.06$ to one significant figure, which is all that we were given in the problem.

(11) X was defined as $[\text{H}_3\text{O}^+]$ at equilibrium but we also need to find pH:

$$\begin{aligned} \text{pH} &= -\log [\text{H}_3\text{O}^+] = -\log (1.04 \times 10^{-3}) = -\log 1.04 - \log (10^{-3}) \\ &= -\log 1.04 - (-3) = 3 - \log 1.04 = 3 - 0.0170 \text{ (from log table)} \\ &= 2.983 \text{ or, since we only had one significant figure, even in the concentrations, pH} = 3.0 \end{aligned}$$

- (B) Since no K is given, we should check a table and conclude that HBr is a strong acid. Hence the only equilibrium involved is the self-ionization of water:



We can consider that HBr is really $\text{H}^+ + \text{Br}^-$, and therefore that $[\text{H}^+] = 0.01 \text{ M}$. Then using $K_w = 1.0 \times 10^{-14} = [\text{H}^+][\text{OH}^-]$, we see that $1.0 \times 10^{-14} = 1 \times 10^{-2} [\text{OH}^-]$ or $[\text{OH}^-] = 1.0 \times 10^{-12} \text{ M}$.

- (C) The question implies a K_a for HCN and therefore that HCN is a weak acid.

(1)		HCN	+	H ₂ O		H ₃ O ⁺	+	CN ⁻
(2) I		1.0 M		-----		0		0
C		2.69 x 10 ⁻⁵ M.				0		0
P		0				2.69 x 10 ⁻⁵ M		2.69 x 10 ⁻⁵ M
E		1.0 M				.69 x 10 ⁻⁵ M		2.69 x 10 ⁻⁵ M

- (3) We are given $\text{pH} = 4.57 = -\log [\text{H}^+]$ at equilibrium.

Then $\log [\text{H}^+] = -4.57 = -5 + 0.43$

From a log table, $[\text{H}^+] = 2.69 \times 10^{-5} \text{ M}$ at equilibrium.

1.0 M is the initial concentration of HCN before ionizing.

There is no $[\text{H}^+]$ or $[\text{CN}^-]$ before ionizing, so none is consumed and no [HCN] is produced.

- (4) Entered in the table in (1).

- (5) The $[\text{H}^+]$ column can be completed: $2.69 \times 10^{-5} = 0 + \text{P} - 0 = \text{P}$

- (6) Using the stoichiometry:

$[\text{CN}^-]$ produced simultaneously with $[\text{H}^+]$, $\text{P} = 2.69 \times 10^{-5} \text{ M}$

$[\text{HCN}]$ consumed in producing $[\text{H}^+]$, $\text{C} = 2.69 \times 10^{-5} \text{ M}$.

- (7) Now using $\text{E} = \text{I} + \text{P} - \text{C}$:

For [HCN]: $\text{E} = 1.0 + 0 - 2.69 \times 10^{-5} \text{ M} = 1.0 \text{ M}$

For $[\text{CN}^-]$: $\text{E} = 0 + 2.69 \times 10^{-5} - 0 = 2.69 \times 10^{-5} \text{ M}$

- (8)
$$K_a = \frac{[\text{H}^+][\text{CN}^-]}{[\text{HCN}]}$$

- (9)
$$K_a = \frac{(2.69 \times 10^{-5})^2}{1.0} = 7.24 \times 10^{-10} \quad \text{or} \quad 7.2 \times 10^{-10}$$

- (10)

- (11) Since only K_a was requested, this completes the solution of the problem.

- (D) Since K_b is given, we know that CH_3NH_2 (which can be thought of as NH_3 with one H replaced by a methyl group) must be a weak base.

(1)		CH ₃ NH ₂	+	H ₂ O		CH ₃ NH ₃ ⁺	+	OH ⁻
(2) I		0.2 M		-----		0		0
C		X M.				0		0
P		0				X M.		X M
E		(0.20 - X) M				X M		X M

- and (3) 0.20 M. is $[\text{CH}_3\text{NH}_2]$ before reaction.
- (4) We need to find pOH, so it makes sense to let $X = [\text{OH}^-]$ at equilibrium. There is neither CH_3NH_3^+ nor appreciable OH^- before the reaction so set their initial concentrations as zero, and consequently their C's and the P position of CH_3NH_2 as zero. (Note that we can ignore the H_2O in the table since it is included in K_b .)
- (5) The $[\text{OH}^-]$ column is now ready to apply $E = I + P - C$:
 $X = 0 + P - 0 = P$
- (6) If X moles/liter of OH^- are formed, P for CH_3NH_3^+ must also be X and X moles/liter of CH_3NH_2 must be consumed by the reaction.
- (7) Now for CH_3NH_2 : $E = 0.20 + 0 - X = 0.20 - X$
 and for CH_3NH_3^+ : $E = 0 + X - 0 = X$
- (8)
$$K_b = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]} = 2.3 \times 10^{-10} \quad (\text{given})$$
- (9)
$$2.3 \times 10^{-10} = \frac{X^2}{0.20 - X}$$
- (10) Since $K_b \ll 1$, assume $X \ll 0.20$ and $0.20 - X \approx 0.20$
 Then $2.3 \times 10^{-10} = \frac{X^2}{0.20}$ $X^2 = 4.6 \times 10^{-11}$ $X = 6.8 \times 10^{-6}$
Check approximation: $0.20 - 6.8 \times 10^{-6} \approx 0.20$
- (11) The problem asks for $\text{pOH} = -\log [\text{OH}^-] = -\log (6.8 \times 10^{-6})$
 $= -\log 6.8 - (-6) = 6 - 0.83 = 5.17$, or 5.2.

Common Ion Effect and Buffers

In the problems previously encountered involving weak acids and bases the initial concentrations of the products have always been zero. In 'Common Ion' problems the initial concentration of one of the products is not zero because a salt containing the anion or an extra source of H_3O^+ or OH^- has been added. These problems are worked by the same method as weak acid and base problems except for an extra entry in your table under initial concentration.

A buffer is a mixture of a weak acid and its conjugate weak base. If there are reasonably large amounts of both species present, then the pH will be relatively unaffected when acid or base is added, since the equilibrium will shift to absorb the addition, in accord with Le Chatelier's Principle.

Suggested method:

- (1) Decide what the predominant species will be in the solution, after equilibrium is established. There will be two predominant species in a common ion problem.
- (2) Write the chemical equilibrium equation which involves the species from (1).
- (3) Prepare a table of I, C, P, E as in other equilibrium problems.
(a) Choose X to be the equilibrium concentration of the species in (2) which you decided in (1) would not be predominant.
- (4) Write the equilibrium constant expression for the equation in (2) and find its value.
- (5) Using the value of K and the number of significant figures being used, decide what simplifying assumptions to use. Calculate X.
- (6) Check assumptions.
- (7) Calculate whatever quantities the problem requested.

Examples:

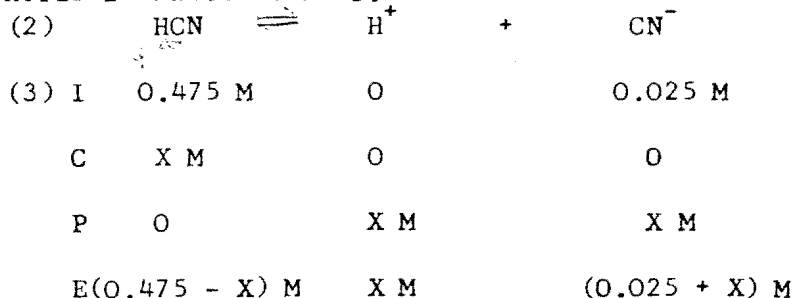
- (A) 500 ml. of 1.0 M HCN is mixed with 500 ml. of 0.05 M NaOH. What is the $[\text{H}_3\text{O}^+]$ of the resulting solution?
- (B) What is the $[\text{H}_3\text{O}^+]$ of a solution prepared by adding 0.75 moles of acetic acid and 0.25 moles of sodium acetate, to enough water to make one liter of solution?
- (C) What is the $[\text{H}_3\text{O}^+]$ of a solution prepared by adding 1.00 moles of acetic acid and 0.25 moles of a strong base to enough water to make one liter of solution?
- (D) Calculate the pH of a solution prepared by mixing equal volumes of 1.50 molar acetic acid and 2.50 molar sodium hydroxide.
- (E) Calculate the $[\text{CN}^-]$ and pH of a solution which is 0.1 M in HCN and 0.1 M in HCl.
- (F) Decide which of the above should be good buffers and calculate the new pH after 0.1 mole of strong acid per liter is added, assuming no change in volume, for examples (A) and (B). What would be the new pH if this amount of strong acid were added to water?

The solutions to these examples are worked out on the following pages.

Solutions:

(A) After combining the two volumes, the concentrations will be 0.5 M HCN and 0.025 M OH⁻, before any acid-base reaction occurs.

(1) A part of the HCN will be neutralized by the OH⁻, so the predominant species will be HCN and CN⁻. Most of the OH⁻ will be used up, and very little H⁺ will be present, since the equilibrium $K_w = [H^+][OH^-] = 1 \times 10^{-14}$ always holds in water at 25°C.



X is defined as the [H⁺] at equilibrium. If we were to let 1.0×10^{-7} be the initial value for [H⁺], we would drop it as insignificant compared to 0.025 and 0.475 in the equilibrium values for HCN and CN⁻, so it would not affect the results.

$$(4) \quad K_a = \frac{[H^+][CN^-]}{[HCN]} = \frac{X(0.025 + X)}{0.475 - X} = 7.2 \times 10^{-10}$$

(5) Since K_a is very small, the numerator must be very small, hence X must be very small. Assume $0.025 + X \approx 0.025$ and $0.475 - X \approx 0.475$.

$$\frac{0.025X}{0.475} = 7.2 \times 10^{-10} \quad \text{so } X = \frac{7.2 \times 10^{-10} \times 4.75 \times 10^{-1}}{2.5 \times 10^{-2}} = 1.4 \times 10^{-8}$$

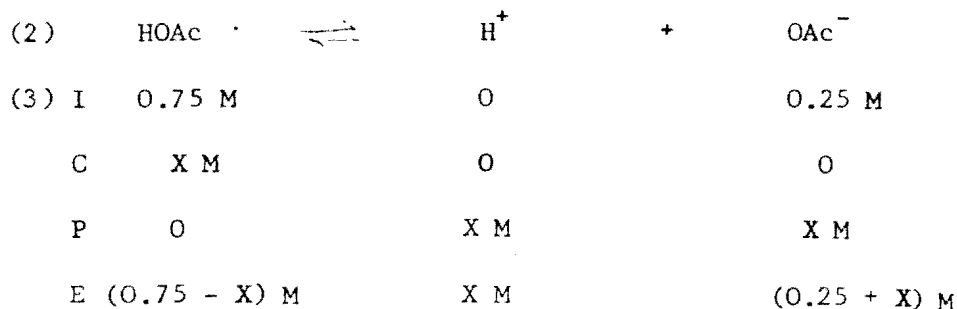
(6) Check assumption: X is beyond the significant figures of 0.025 and 0.475.

(7) Only [H⁺] was requested, and it is X, but we will calculate the pH to compare in Example (F).

$$\text{pH} = 8 - \log 1.4 = 8 - .14 = 7.9$$

(B) 0.75 M HOAc, 0.25 M OAc⁻

(1) There is no neutralization involved, so these will be the predominant species at equilibrium and the H⁺ concentration will not be very large.



Here X was defined as the equilibrium concentration of H⁺ and the rest of the entries were derived from that.

$$(4) K_a = \frac{[H^+][OAc^-]}{[HOAc]} = 1.8 \times 10^{-5} = \frac{X(0.25 + X)}{0.75 - X}$$

(5) K_a is not as small as one might wish, but assume X is small enough to let $0.25 + X \approx 0.25$ and $0.75 - X \approx 0.75$. Be sure to check when we find X .

$$\frac{0.25 X}{0.75} = 1.8 \times 10^{-5} \quad \text{so } X = 5.4 \times 10^{-5}$$

(6) Check: $0.25 + 0.000054 \approx 0.25$ and $0.75 - 0.000054 \approx 0.75$

(7) Calculate: $pH = 5 - \log 5.4 = 5 - .73 = 4.27$ for comparison in Example (F)

(C) 1.00 M HOAc, 0.25 M OH^-

(1) Approximately 0.25 M of the HOAc will be neutralized to form OAc^- , so the predominant species will be HOAc and OAc^- . There will not be a large amount of either H^+ or OH^- present.



(3) I 0.75 M 0 0.25 M

C

P

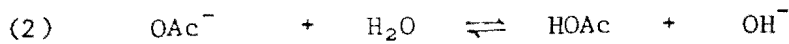
E

X M

At this point we see that this problem is identical to (B), so see that solution.

(D) When the volumes are combined, the concentrations become 0.75 M HOAc and 1.25 M OH^- .

(1) Here we have an excess of OH^- , so essentially all of the HOAc will be converted to OAc^- and an excess of OH^- will also be present. Very little HOAc will be present at equilibrium.



(3) I 0.75 M ---- 0 0.50 M

C X M ---- 0 0

P 0 X M X M.

E (0.75 - X) M X M (0.50 + X) M

X is defined as $[HOAc]$ at equilibrium.

$$(4) K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = \frac{[HOAc][OH^-]}{[OAc^-]} = \frac{X(0.50 + X)}{0.75 - X} = 5.56 \times 10^{-10}$$

(5) Assume $0.50 + X \approx 0.50$ and $0.75 - X \approx 1.25$

$$\frac{0.50 X}{0.75} = 5.56 \times 10^{-10} \quad \text{or } X = 8.34 \times 10^{-9}$$

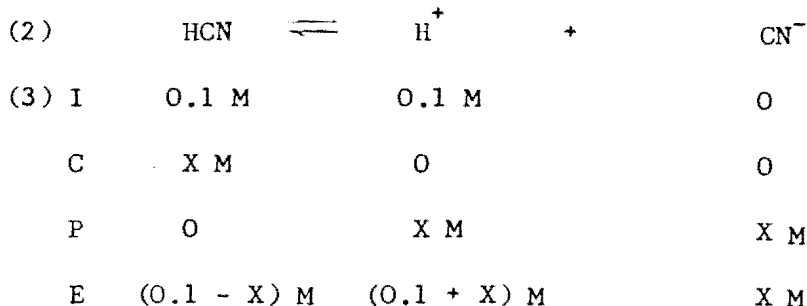
(6) This is certainly small enough to neglect.

(7) Calculate the pH: $[\text{OH}^-] = 0.50 \text{ M}$ so $[\text{H}^+] = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{5 \times 10^{-1}} = 2 \times 10^{-14}$

$$\text{and pH} = 14 - \log 2 = 14 - 0.3 = 13.7$$

(E) $[\text{HCN}] = 0.1 \text{ M}$ and $[\text{H}^+] = 0.1 \text{ M}$

(1) The presence of the excess H^+ will repress the ionization of the HCN, according to Le Chatelier's Principle, so the predominant species will be HCN and H^+ . There will be very little CN^- present.



$$(4) \quad K_a = \frac{[\text{H}^+][\text{CN}^-]}{[\text{HCN}]} = 7.2 \times 10^{-10} = \frac{X(0.1 + X)}{0.1 - X}$$

(5) Either the numerator is small or the denominator is large to give such a small constant, and either would be effected by having X very small. So assume $0.1 + X \approx 0.1$ and $0.1 - X \approx 0.1$

$$\text{then } \frac{X(0.1)}{0.1} = 7.2 \times 10^{-10} \text{ or } X = 7.2 \times 10^{-10}$$

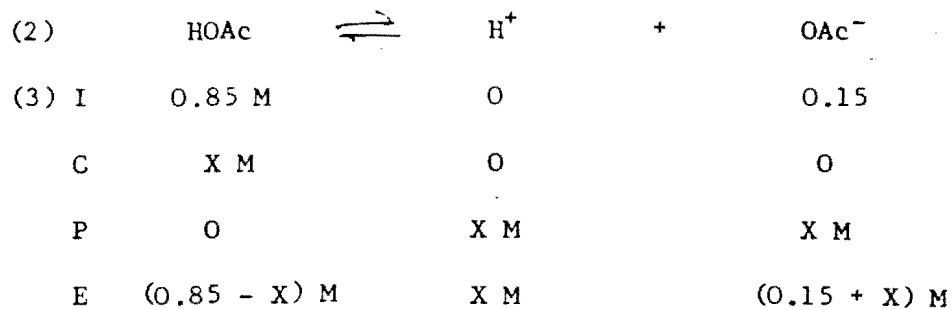
(6) This confirms our approximation.

(7) $[\text{H}^+] = 0.1$ so $\text{pH} = -\log 1 \times 10^{-1} = 1$.

(F) Since the amount of one of the conjugate species is very small in examples (D) and (E) we know it would be used up or increased very easily by addition of acid or base, so (D) and (E) would not be good buffers. (B) should be a better buffer than (A), since the amounts of the conjugate species are more nearly equal in (B). We will check this in our calculations:

(A)(1) With the 0.1 M acid, essentially all the 0.025 M OH^- will be neutralized, and the predominant species will be HCN and H^+ . Hence we will not have enough CN^- remaining to serve as a buffer. The $[\text{H}^+] = 0.075$, $\text{pH} = 1.12$. It was 7.9.

(B)(1) The 0.1 moles of acid will convert OAc^- to HOAc but there will still be 0.15 M OAc^- left and the predominant species will be HOAc and OAc^- and there will be very small quantities of either H^+ or OH^- .



(4)
$$K_a = \frac{[H^+][OAc^-]}{[HOAc]} = 1.8 \times 10^{-5} = \frac{X(0.15 + X)}{0.85 - X}$$

(5) Assume X is small so: $0.15 + X \approx 0.15$ and $0.85 - X \approx 0.85$

$$\frac{0.15 X}{0.85} = 1.8 \times 10^{-5} \text{ so } X = 1.02 \times 10^{-4}$$

(6) $0.15 + 0.0001 \approx 0.15$ and $0.85 - 0.0001 \approx 0.85$

(7) Calculate: $pH = 4 - \log 1.02 = 4.$

So we see that the pH has only changed from 4.27 to 4 when the 0.1 mole per liter of acid was added.

If 0.1 mole per liter of acid were added to water the pH would be:

$$pH = -\log 1 \times 10^{-1} = 1$$

Hence (C) is a good buffer.