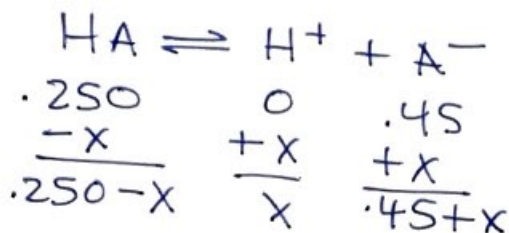


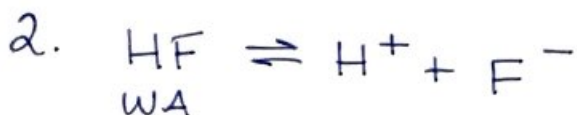
Acids, Bases, Buffers Practice Questions (Fri. 3/26)



$$1.0 \times 10^{-6} = \frac{(x)(.45+x)}{(.250-x)} \quad \begin{array}{l} \text{5\%} \\ \text{Rule} \end{array}$$

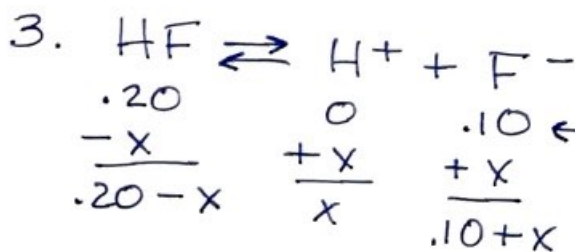
$$x = [\text{H}^+] = 5.6 \times 10^{-7} \text{ M}$$

$$\boxed{\text{pH} = 6.25}$$



from the salt

The solution would contain significant amounts of both H^+ and $\text{F}^- \rightarrow$ a Buffer!!
 If HCl is added to the solution, it will be neutralized by the F^- (SCB).



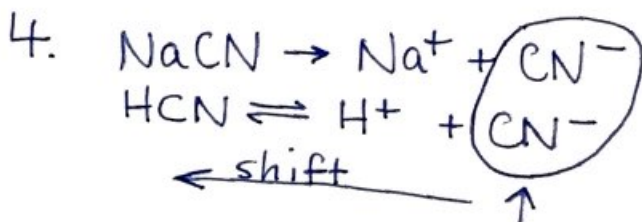
Look this up!
 I forgot :)

$$K_a \text{ of } \text{H}_f = 3.2 \times 10^{-4}$$

$$3.2 \times 10^{-4} = \frac{(x)(.10)}{(.20)} \quad \begin{array}{l} \text{5\%} \\ \text{Rule} \end{array}$$

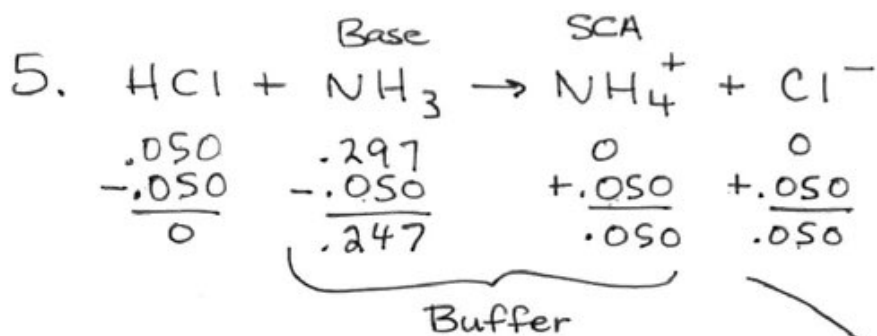
$$x = [\text{H}^+] = 6.4 \times 10^{-4} \text{ M}$$

$$\boxed{\text{pH} = 3.19}$$



The common ion will shift the acid equilibrium to the left, causing the $[\text{H}^+]$ to decrease and pH to increase. Bis False

** CHOICE E. If strong base is added, it will be neutralized by the
 $\text{HCN} + \text{OH}^- \rightarrow \text{CN}^- + \text{H}_2\text{O}$

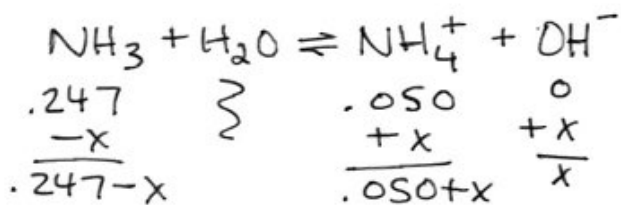


$$\begin{array}{l}
 \text{HCl} \\
 (100.0)(.100) = M_2(200.0) \\
 M_2 = .050 \text{ M} \\
 \text{NH}_3 \\
 (100.0)(.594) = M_2(200.0) \\
 M_2 = .247 \text{ M}
 \end{array}$$

$$\text{pH} = \text{pK}_a + \log \frac{\text{B}}{\text{A}} \\
 \text{pH} = -\log \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} + \log \frac{.247}{.050}$$

$$\boxed{\text{pH} = 9.95}$$

OR work the neutralization and then find the new equilibrium

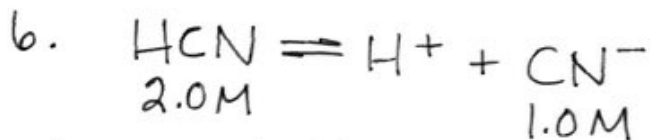


$$K_b = 1.8 \times 10^{-5} = \frac{(.050)(x)}{(.247)}$$

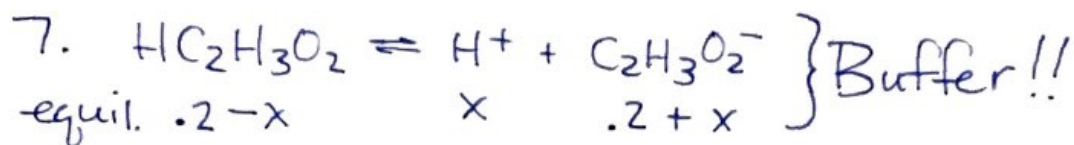
$$x = [\text{OH}^-] = 8.9 \times 10^{-5}$$

$$\text{pOH} = 4.04$$

$$\boxed{\text{pH} = 9.95}$$



- F A. The solution is a buffer because it contains a weak acid and its conjugate base.
- F B. and C. When a buffer is made with a weak acid, its pH will be < 7 and $[\text{OH}^-] < [\text{H}^+]$
- F D. This buffer contains more acid than base, so will be able to neutralize more strong base than strong acid.
- * E. All are false!



equil. $.2 - x$

x

$.2 + x$

$$1.8 \times 10^{-5} = \frac{(x)(.2+x)}{(.2-x)}$$

$$x = 1.8 \times 10^{-5} \text{ M}$$

$$\text{pH} = 4.74$$

OR

$$\text{pH} = \text{pK}_a + \log \frac{B}{A}$$

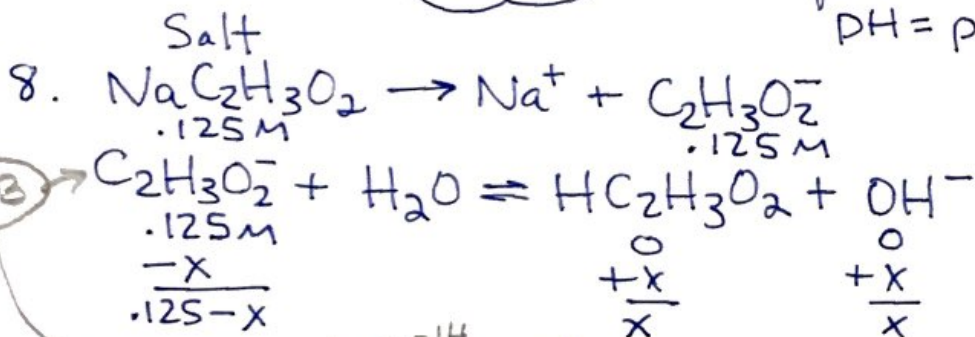
$$= -\log 1.8 \times 10^{-5} + \log \frac{.2}{.2}$$

$$= 4.74$$

$$\log 1 = 0$$

Handy Fact

* when you have a buffer with equimolar acid and base, $\text{pH} = \text{pK}_a$



salts can produce ions that go into hydrolysis

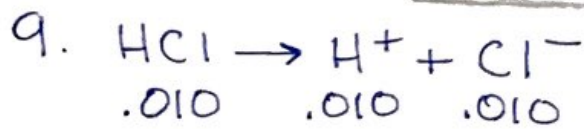
SCB → need K_b

$$\frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = \frac{x^2}{.125}$$

$$x = [\text{OH}^-] = 8.3 \times 10^{-6} \text{ M}$$

$$\text{pOH} = 5.08$$

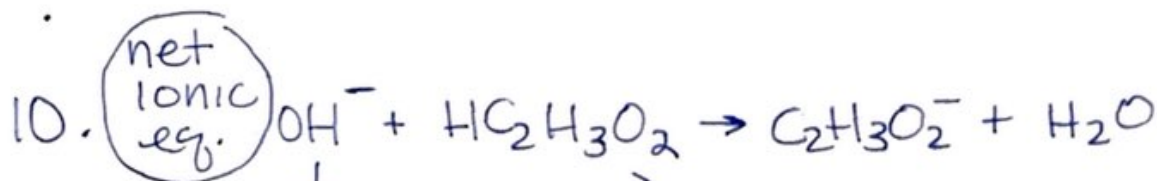
$$\text{pH} = 8.92$$



$$\text{pH} = -\log .010$$

$$\text{pH} = 2.00$$

$[\text{H}^+]$ produced by weak acid is negligible.

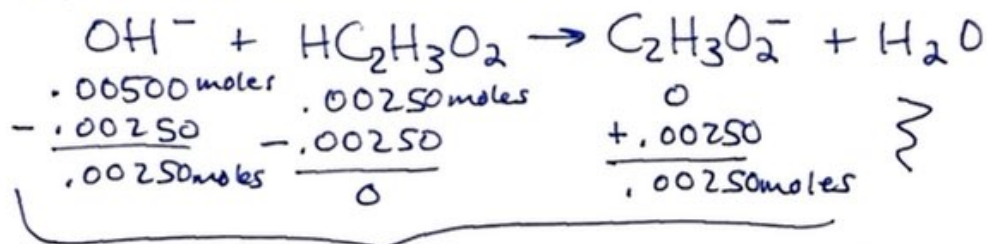


Just do this in calc

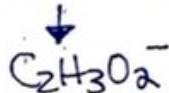
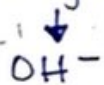
$$.0500 \text{ L} \times \frac{.100 \text{ moles}}{1 \text{ L}} = .00500 \text{ moles}$$

$$.0250 \text{ L} \times \frac{.100 \text{ moles}}{1 \text{ L}} = .00250 \text{ moles}$$

neutralization reaction:



At the end of the neutralization reaction, the solution contains strong base and weak base.



pH is determined by strong base:

$$\text{pOH} = -\log \frac{.00250 \text{ moles}}{.0750 \text{ L}}$$

$$= 1.477$$

$$\boxed{\text{pH} = 12.523}$$

11. To make a buffer with a pH of 9.26, should use a weak base and its conjugate. So C or D

D is BEST because it is the most concentrated and has a greater capacity to neutralize strong acid or strong base.

* [Buffer capacity] *

↓
depends on concentration