Chapter 13 AP Chem
Acid/Base Review

1. Strong acid/base: Easy Peasy☺!

A. Strong acids 100% ionize in water, so $[H^+]=[H_3O^+]=[HA]$.
0.25 M HCl has 0.25 M $[H^+]$=0.25 and pH = -log$[H^+]$= -log(0.25) = 0.60 = pH

B. Strong bases 100% dissociate (already ionic)
0.25 M NaOH, $[OH^-]$ = 0.25 and pOH= - log(OH) = - log(0.25) = 0.60, pH = 14 - pOH=13.40 = pH

0.25 M $Ca(OH)_2$, remember $[OH^-]$ = 2x = 0.50, pOH=0.30 and pH = 13.70

2. Weak acid/base, little more challenging…require a K_a or K_b expression. Weak acids and bases ionize water way less than 100% (about 1% or less, maybe up to 5% for the best ionizers like HF)

HF $K_a = 7.2x10^{-4}$, CH_3COOH $K_a = 1.8x10^{-5}$ NH_3 $K_b=1.8x10^{-5}$

A. Weak acids in general, $HA + H_2O \leftrightarrow H_3O^+ + A^-$ and $K_a = \frac{[H_3O^+][A^-]}{[HA]}$

So 0.25 M CH_3COOH
First, do a RICE table for the equilibrium:

R	HA +	$H_2O \leftrightarrow$	H_3O^+ +	A^-
I	0.25		0	0
C	-x		+x	+x
E	0.25-x		x	x

and $K_a = \frac{[H_3O^+][A^-]}{[HA]} = \frac{[x][x]}{[0.25-x]} = 1.8x10^{-5}$, and x = 0.00212
so pH=-log$[H_3O^+]$ = -log[0.00212] = 2.7 = pH for 0.25 M CH_3COOH
which is not as acidic as strong acid (0.25M strong HCl has pH = 0.6)

B. Weak base in general $B + H_2O \leftrightarrow HB^+ + OH^-$ and $K_b = \frac{[HB^+][OH^-]}{[B]}$

So 0.25 M NH_3
First, do a RICE table for the equilibrium:

R	NH_3 +	$H_2O \leftrightarrow$	NH_4^+ +	OH^-
I	0.25		0	0
C	-x		+x	+x
E	0.25-x		x	x

and $K_b = \frac{[HB^+][OH^-]}{[B]} = \frac{[x][x]}{[0.25-x]} = 1.8x10^{-5}$ so x = 0.00212

and pOH = -log(0.00212) = 2.7, and pH = 14-pOH = 14-2.7= 11.3 = pH for 0.25 M

Next page, salts of weak acids/bases and buffers, half neutralizations….

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3. Salt of weak acid/base. How do you know?

NaCl is the product of a strong acid/strong base, so it is neutral (HCl + NaOH → NaCl + H_2O).
If you have a weak acid/strong base, there will be a conjugate base!
For example, HF + NaOH → NaF + H_2O but F^- is the "strong" conjugate base of a weak acid!
Now what?!? $K_a K_b = K_w$ so for HF with $K_a=7.2x10^{-4}$, the K_b of F^- is K_w / K_a then do RICE table:

A. 0.25 M NaF will be basic, but not as basic as 0.25 M NaOH (pH=13.4 STRONG base)
and not as basic as 0.25 M NH_3 (pH=11.3 weak base)
K_b of F^- is $K_w / K_a = 1.0x10^{-14} / 7.2x10^{-4} = 1.39x10^{-11}$
$F^- + H_2O \leftrightarrow HF + OH^-$ and $K_b = [HF][OH^-] / [F^-] = x^2 / (0.25-x) = 1.39x10^{-11}$
Solved $x = 1.86x10^{-6} = [OH]$ so pOH = 5.7 and <u>pH = 8.3</u> (compared to 13.4 and 11.3)

4. Buffer!

Mix a weak acid and a salt (of that acid) or a weak base/salt. This time the RICE table 'I' values are not zero.

0.25 M HF and 0.1 M NaF. A buffer is a mixture of acid/salt.

R	HF +	$H_2O \leftrightarrow$	$H_3O^+ +$	F^-
I	0.25		0	0.1
C	-x		+x	0.1+x
E	0.25-x		x	0.1+x
approx.	0.25		x	0.1

And since X is small, use the approximations so $K_a = 0.1x / 0.25$ (too bad we can't use solver on UIL, but you can on AP!)

$x=K_a*0.25/0.1 = 1.8x10^{-3} = H_3O^+$ so pH = $-\log(1.8x10^{-3})$ = <u>pH= 2.7</u>

5. Half-titration of weak acid / strong base (or strong acid/weak base).

At the halfway point, the acid is half gone and the salt is half produced--so it is a buffer!
So for 50 mL of 0.25 M HF and 25 mL of 0.25 M NaOH, half the acid is neutralized, making half of the salt (NaF). So 6.25 mmol HF left and 6.25 mmol NaF produced in 75 mL, so $[HF]=[F^-]=0.0833$

$K_a = 0.0833 x / 0.0833$ (who cares what the $[HF]=[F^-]$ could be, they cancel!)

So Ka = x and pH = pKa, so pH = $-\log(7.2x10^{-4})$ = <u>pH =3.14</u> (compared to #4 pH=2.7)