Today 2/12 Sections 15.8-15.10, 15.13

$[H_3O^+][OH^-] = K_w = 1\times10^{-14}$

$[H_3O^+] \longleftrightarrow [OH^-]$

$pH = -\log[H_3O^+]$ | $[H_3O^+] = 10^{-pH}$

$pOH = -\log[OH^-]$ | $[OH^-] = 10^{-pOH}$

$pH \longleftrightarrow pOH$

$pH + pOH = 14$

universal indicator

pH	
4	red
5	orange
6	yellow
7	green
8	~~purple~~ blue
9	blue
10	violet

- strong acids dissociate 100%:

$HCl(aq) + H_2O(l) \xrightarrow{100\%} H_3O^+ + Cl^-$

- weak acids:

$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$

$K_a^{HA} = \dfrac{[H_3O^+][A^-]}{[HA]}$

$pK_a^{HA} = -\log K_a$

$K_a = 10^{-pK_a}$

strong base:

$NaOH(aq) \xrightarrow{100\%} OH^- + Na^+$ no K (because no equilibrium)

weak base:

$A^- + H_2O \rightleftharpoons OH^- + HA$ $K_b = \dfrac{[OH^-][HA]}{[A^-]}$ $pK_b = -\log K_b$ $K_b = 10^{-pK_b}$

$2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$ $K_w = 1\times10^{-14} = [H_3O^+][OH^-]$

$pK_w = 14$

- what is the pH of a 6.7×10^{-4} M $HNO_3(aq)$ ~~sto~~ solution?

$HNO_3 + H_2O \xrightarrow{100\%} H_3O^+ + NO_3^-$

$[H_3O^+] = \underline{6.7}\times10^{-4}$ M (two sig figs)

$pH = -\log(6.7\times10^{-4}) = 3.\underline{17}$ (two sig figs)

- What is the pH of a 0.42 M $HC_2H_3O_2(aq)$ solution?

M	$HC_2H_3O_2$	+ H_2O $\rightleftharpoons$	H_3O^+	+ $C_2H_3O_2^-$
I	0.42		0	0
C	$-x$		$+x$	$+x$
E	$0.42-x \approx 0.42$		$x = 2.73\times10^{-3} = x$	

$K_a = 1.78\times10^{-5}$

changes will always be 1x for acid/base chemistry

$$K_a = 1.78\times10^{-5} = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} = \frac{x^2}{0.42-x}$$

"400 Rule": Is the $\frac{[HA]_I}{K_a^{HA}} > 400$? $\frac{0.42}{1.78\times10^{-5}} = 2.4\times10^4$ ✓

if yes, x is very small. If no, use quadratic

so...

$$\frac{x^2}{0.42} = 1.78\times10^{-5} \qquad x = 2.73\times10^{-3}$$

$[H_3O^+] = 2.73\times10^{-3}$

$pH = -\log(2.73\times10^{-3})$

$pH = 2.56$

% dissociation $= 100\% * \frac{x}{[HA]_I} = \frac{2.73\times10^{-3}}{.42}\times100 = 0.65\%$

· An unknown acid gives a pH = 3.11 for a 0.150 M solution. What is its K_a? and pK_a?

M	HA	+ H_2O ⇌	H_3O^+	+ A^-
I	0.150		0	0
C	$-x$		$+x$	$+x$
E	$0.150-x$		x	x

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} = \frac{x^2}{0.15-x}$$

$$K_a = \frac{(7.76\times10^{-4})^2}{(0.150-7.76\times10^{-4})}$$

$$\boxed{K_a = 4.02\times10^{-6}}$$

$pH = 3.11$

$[H_3O^+] = 10^{-3.11} = 7.76\times10^{-4}\ M = x$

$$\boxed{pK_a = -\log(4.02\times10^{-6}) = 5.40}$$

· What is the pH of a 0.020 M HNO_2 solution given $K_a = 4.5\times10^{-4}$?

M	HNO_2	+ H_2O ⇌	H_3O^+	+ NO_2^-
I	0.020		0	0
C	$-x$		$+x$	$+x$
E	$0.020-x$		x	x

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = 4.5\times10^{-4}$$

$$\frac{x^2}{0.20-x} = 4.5\times10^{-4}$$

$$x^2 = (4.5\times10^{-4})(.02-x)$$

400 Rule: $\frac{0.020}{4.5\times10^{-4}} = 44$ X (use quadratic)

$$x^2 = 9.0\times10^{-6} - 4.5\times10^{-4}x$$

$$\underbrace{x^2}_{a} + \underbrace{4.5\times10^{-4}x}_{b} \underbrace{- 9.0\times10^{-6}}_{c} = 0$$

$x = 2.78\times10^{-3}$ and -3.23×10^{-3}

pick positive

$[H_3O^+] = 2.78\times10^{-3}\ M$

$pH = -\log(2.78\times10^{-3})$

$pH = 2.56$

% dissociation = $100\% \times \frac{2.78\times10^{-3}}{0.020}$ = 14%

· What is the pH of a 1.7×10^{-3} M KOH solution?

$[OH^-] = 1.7\times10^{-3}$ M

pOH = $-\log 1.7\times10^{-3}$ M = 2.77

pH: 14 - 2.77 = 11.23

Suppose the solution was then diluted 25mL up to 100mL with H_2O. What is the new pH?

Dilution formula: $M_cV_c = M_dV_d$

$(1.7\times10^{-3})(25mL) = (100mL)(M_d)$

$M_d = 4.25\times10^{-4}$ M = $[OH^-]$

pOH = $-\log 4.25\times10^{-4}$ M = 3.37

$\boxed{pH = 14 - 3.37 = 10.63}$

* dilutions "force things to the middle" where pH and pOH will get closer to 7

· What is the pH of a 0.37M NaF solution? weak base = F^- weak acid = HF

(Na^+, F^-)

wa	wb
HA	A^-

$K_a = \frac{[H_3O][A^-]}{[HA]}$ $K_b = \frac{[OH^-][HA]}{[A^-]}$

$\boxed{K_w = K_a * K_b}$ $\frac{[H_3O][A^-][OH^-][HA]}{[HA][A^-]} = [H_3O^+][OH^-] = K_w$

· Given: $K_a = 3.5\times10^{-4}$

M	F^- + H_2O ⇌	OH^-	+ HF
I	0.37	0	0
C	-x	+x	+x
E	0.37-x	x	x

$K_b^{F^-} = \frac{[OH^-][HF]}{[F^-]}$

$2.86\times10^{-11} = \frac{x^2}{.37}$

$x = 3.25\times10^{-6} = [OH^-]$

$K_b = \frac{K_w}{K_a} = \frac{1.0\times10^{-14}}{3.5\times10^{-4}} = 2.86\times10^{-11}$

400 rule - $\frac{0.37}{2.86\times10^{-11}} > 400$? yes ✓

pOH = $-\log(3.25\times10^{-6}$ M$)$

pOH = 5.49

$\boxed{pH = 8.51}$