

Today 2/12 Sections 15.8-15.10, 15.13

$$[H_3O^+][OH^-] = K_w = 1 \times 10^{-14}$$



$$pH = -\log [H_3O^+]$$

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

$$pOH = -\log [OH^-]$$

$$[OH^-] = 10^{-pOH}$$

$$pH \longleftrightarrow pOH$$

$$pH + pOH = 14$$

universal indicator

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

strong acids dissociate 100%:



weak acids:

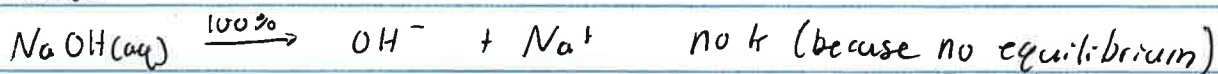


$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

$$pK_a = -\log K_a$$

$$K_a = 10^{-pK_a}$$

strong base:

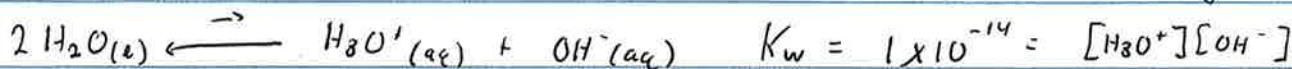


weak base:



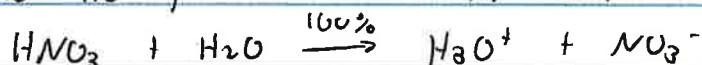
$$K_b = \frac{[OH^-][HA]}{[A^-]} \quad pK_b = -\log K_b$$

$$K_b = 10^{-pK_b}$$



$$pK_w = 14$$

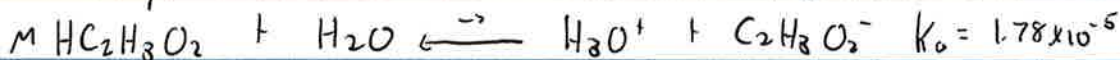
What is the pH of a $6.7 \times 10^{-4} M$ $HNO_3(aq)$ solution?



$$[H_3O^+] = 6.7 \times 10^{-4} M \text{ (two sig figs)}$$

$$pH = -\log(6.7 \times 10^{-4}) = 3.17 \text{ (two sig figs)}$$

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



I	0.42	0	0
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C	-x	+x	+x
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E	$0.42 - x = 0.42$	$x = 2.73 \times 10^{-3} = x$	
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changes will always be x for acid/base chemistry

~~What is the pH of a 0.42 M HC2H3O2(aq) solution?~~

$$K_a = 1.78 \times 10^{-5} = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} = \frac{x^2}{0.42 - x}$$

"400 Rule": Is the $\frac{[HA]_I}{K_a} > 400$? $\frac{0.42}{1.78 \times 10^{-5}} = 2.4 \times 10^4$ ✓

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

So...

$$\frac{x^2}{0.42} = 1.78 \times 10^{-5} \quad x = 2.73 \times 10^{-3}$$

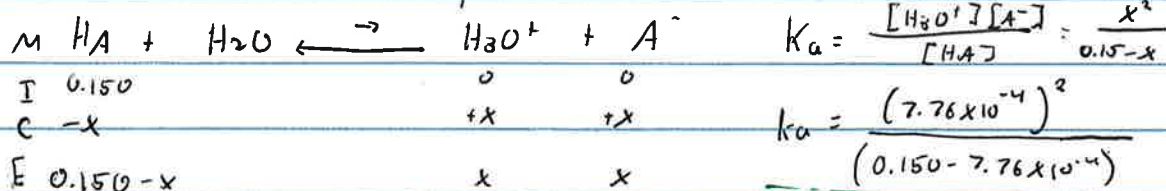
$$[H_3O^+] = 2.73 \times 10^{-3}$$

$$pH = -\log(2.73 \times 10^{-3})$$

$$pH = 2.56$$

$$\% \text{ dissociation} = 100\% * \frac{x}{[HA]_I} = \frac{2.73 \times 10^{-3}}{.42} \times 100 = 0.65\%$$

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



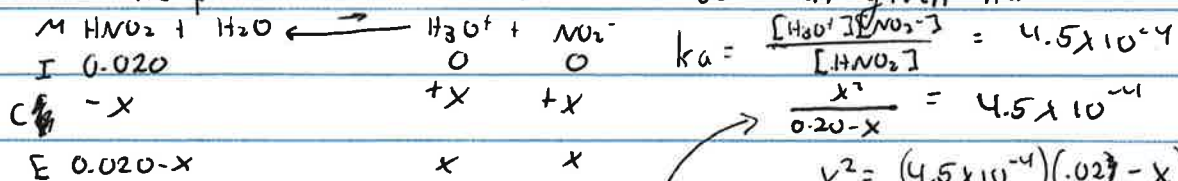
$$pH = 3.11$$

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

$$K_a = 4.02 \times 10^{-6}$$

$$pK_a = -\log(4.02 \times 10^{-6}) = 5.40$$

What is the pH of a $0.020 M$ HNO_2 solution given $K_a = 4.5 \times 10^{-4}$?



400 Rule: $\frac{0.020}{4.5 \times 10^{-4}} = 44$ ✗ (use quadratic) $x^2 = 9.0 \times 10^{-6} - 4.5 \times 10^{-4}x$

$$[H_3O^+] = 2.78 \times 10^{-3} M$$

$$pH = -\log(2.78 \times 10^{-3})$$

$$pH = 2.56$$

$$x^2 + 4.5 \times 10^{-4}x - 9.0 \times 10^{-6} = 0$$

$$x = 2.78 \times 10^{-3} \text{ and } -3.23 \times 10^{-3}$$

pick positive

$$\% \text{ dissociation} = 100\% \times \frac{2.78 \times 10^{-3}}{0.020} = 14\%$$

What is the pH of a $1.7 \times 10^{-3} \text{ M KOH}$ solution?

$$[\text{OH}^-] = 1.7 \times 10^{-3} \text{ M}$$

$$\text{pOH} = -\log 1.7 \times 10^{-3} \text{ M} = 2.77$$

$$\text{pH} = 14 - 2.77 = 11.23$$

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

Dilution formula: $M_c V_c = M_d V_d$

$$(1.7 \times 10^{-3})(25 \text{ mL}) = (100 \text{ mL})(M_d)$$

$$M_d = 4.25 \times 10^{-4} \text{ M} = [\text{OH}^-]$$

$$\text{pOH} = -\log 4.25 \times 10^{-4} \text{ M} = 3.37$$

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

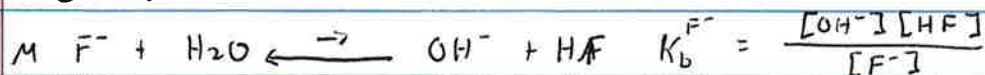
* dilutions "force things to the middle" where pH and pOH will get closer to 7

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

$$\frac{w_a}{HA} \quad \frac{w_b}{A^-} \quad K_a = \frac{[\text{H}_3\text{O}^+][A^-]}{[HA]} \quad K_b = \frac{[\text{OH}^-][HA]}{[A^-]}$$

$$K_w = K_a * K_b \quad \frac{[\text{H}_3\text{O}^+][A^-][\text{OH}^-][HA]}{[HA][A^-]} = [\text{H}_3\text{O}^+][\text{OH}^-] = K_w$$

Given: $K_a = 3.5 \times 10^{-4}$



$$\text{I } 0.37$$

$$\text{C } -x$$

$$\text{E } 0.37 - x$$

$$0 \quad 0$$

$$+x \quad +x$$

$$x \quad x$$

$$2.86 \times 10^{-11} = \frac{x^2}{.37}$$

$$x = 3.25 \times 10^{-6} = [\text{OH}^-]$$

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{3.5 \times 10^{-4}} = 2.86 \times 10^{-11}$$

$$400 \text{ rule} = \frac{0.37}{2.86 \times 10^{-11}} > 400? \text{ yes } \checkmark$$

$$\text{pOH} = -\log(3.25 \times 10^{-6} \text{ M})$$

$$\text{pOH} = 5.49$$

$$\text{pH} = 8.51$$