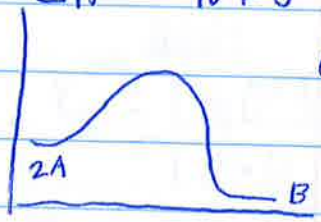
Chapter 16! Day 1
sections 16.1 - 16.7

Warm-up: (from previous worksheet)

$2A(g) \rightleftharpoons B(g)$ $\Delta H = -10\ kJ$



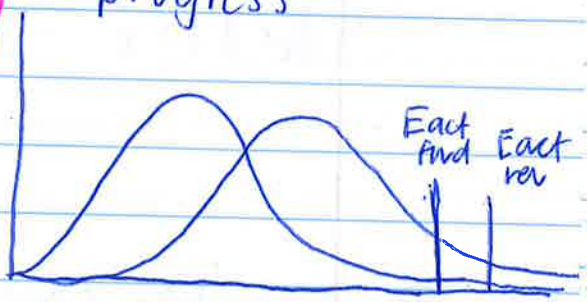
exothermic!

* catalyst speeds up kinetic region but not overall ~~equilibrium~~ K_c

* changing temperature always changes K_c

* changing volume does NOT change K_c

* BUT when volume is changed, moles move left / right depending on increasing / decreasing it



$k = Ae^{-Eact/RT}$

* rate constant only changes when temp. changes

take-aways from 1d-1p on previous worksheet

New Topic: Acid-Base Equilibria

	$2H_2O(l) \rightleftharpoons$	$H_3O^+(aq)$	+	$OH^-(aq)$	
I		0		0	*ignore liquids/ solids
C		+x		+x	
E		x		x	$K_C^{298} = [H_3O^+][OH^-] =$
		$[H_3O^+] =$		$[OH^-] =$	1.0×10^{-14}
		1.0×10^{-7} M		1.0×10^{-7} M	$x^2 = 1.0 \times 10^{-14}$
					$x = 1.0 \times 10^{-7}$

$K_C \rightarrow$ change to K_W

Strong acids (MEMORIZE!)

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

$HBr(aq)$ " "

$HI(aq)$ " "

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

$HClO_4(aq)$ " "

$H_2SO_4(aq)$ " "

* all are strong electrolytes, all dissociate 100%
* other acids not on list = weak acids

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

ex) if $[H_3O^+] = 1.0 \times 10^{-7}$ M plug into calc

$pH = 7.00$ → 2 sig figs, no units

*everything to the right of the decimal = # of sig figs

$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$

use K_w

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

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

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

$pH \longleftrightarrow pOH$

$pH + pOH = 14.00$

What is the pH of the 2.41×10^{-3} M $HClO_4$ soln?

$HClO_4 (aq) + H_2O (l) \xrightarrow{100\%} H_3O^+ + ClO_4^-$

	$HClO_4$		H_3O^+	ClO_4^-
I	2.41×10^{-3}		0	0
C	-2.41×10^{-3}		$x = 2.41 \times 10^{-3}$	$x = 2.41 \times 10^{-3}$
E	0		2.41×10^{-3}	2.41×10^{-3}

$[H_3O^+] = 2.41 \times 10^{-3}$

$pH = -\log (2.41 \times 10^{-3}$

$= 2.618$ need 3 sig figs

becomes exponent, not overall sig fig

$K_c = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = 4.5 \times 10^{-4}$

weak acids

What is pH of .50 M HNO_2 solution?

M	$HNO_2(aq) + H_2O(l) \rightleftharpoons$	$H_3O^+ + NO_2^-$
I	.50 M	0 0
C	$-x$	$+x \quad +x$
E	$.50 - x$	$x \quad x$

$.50 - .01477 = .4852$

$K_a = \frac{x^2}{.50 - x} = 4.5 \times 10^{-4}$

*weak acids barely break apart into ions

*use quadratic

$x^2 + 4.5 \times 10^{-4} x - 2.25 \times 10^{-4} = 0$

$a = 1 \quad b = 4.5 \times 10^{-4} \quad c = -2.25 \times 10^{-4}$

*X → must be positive

$$= \frac{-4.5 * 10^{-4} + \sqrt{(4.5 \times 10^{-4})^2 + (4 \cdot 1 \cdot +2.25 \times 10^{-4})}}{2}$$

$$= \frac{-4.5 \times 10^{-4} + .03000}{2}$$

$x = .0148$

$K_c = (1.477 \times 10^{-2})^2 = 4.5 \times 10^{-4}$

→ .4852

$pH = -\log(1.477 \times 10^{-2})$

$pH = 1.83$

*weak acid equilibrium label K_c as K_a

Tip to avoid quadratic:

Is $.50 - x \approx .50$?

"400 Rule"

IF $\frac{[HA]_I}{K_a} > 400$

is $\frac{.50}{4.5 \times 10^{-4}} > 400$

$1100 > 400 \rightarrow$ YES!

how we can continue... $\frac{x^2}{.50} = 4.5 \times 10^{-4}$

$x = .015$