

6/17/20

11

trial	initial rate	initial [A]	initial [B]	initial [C]
1	6.0×10^{-6}	0.024	0.100	0.032
2	9.6×10^{-5}	0.096	0.100	0.032
3	2.0×10^{-6}	0.024	0.034	0.032
4	4.5×10^{-6}	0.012	0.300	0.064

same

$$R = k [A]^a [B]^b [C]^c$$

$$\frac{R_2}{R_1} = \left(\frac{[A]_2}{[A]_1} \right)^a$$

$$\frac{9.6 \times 10^{-5}}{6.0 \times 10^{-6}} = \left(\frac{0.096}{0.024} \right)^a$$

$$16 = (4)^a \Rightarrow a = 2$$

compare trials 3 + 1

$$\frac{6.0 \times 10^{-6}}{2.0 \times 10^{-6}} = 3 = \left(\frac{0.100}{0.034} \right)^b = 3^b$$

$$b = 1$$

trial	initial rate	initial [A]	initial [B]	initial [C]
1	6.0×10^{-6}	0.024	0.100	0.032
2	9.6×10^{-5}	0.096	0.100	0.032
3	2.0×10^{-6}	0.024	0.034	0.032
4	4.5×10^{-6}	0.012	0.300	0.064

same

trial 4 vs trial 3

$$\frac{4.5 \times 10^{-6}}{2.0 \times 10^{-6}} = \left(\frac{[A]_4}{[A]_3} \right)^2 \left(\frac{[B]_4}{[B]_3} \right)^1 \left(\frac{[C]_4}{[C]_3} \right)^c$$

$$2.25 = \left(\frac{0.012}{0.024} \right)^2 \left(\frac{0.300}{0.034} \right)^1 \left(\frac{0.064}{0.032} \right)^c$$

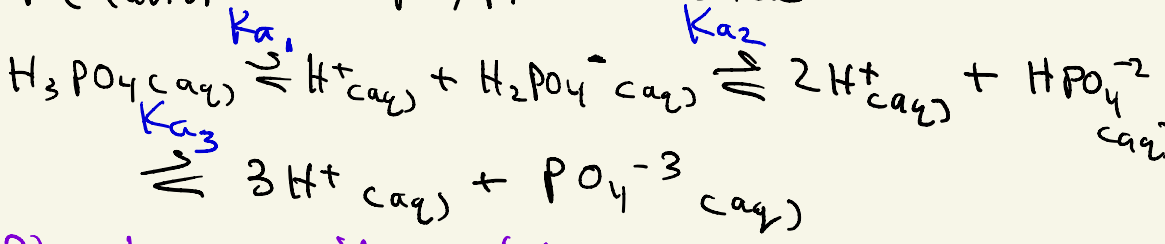
$$\frac{9}{4} = \left(\frac{1}{2} \right)^2 (9)^1 (2)^c$$

$$\left(\frac{9}{4} \right) = \left(\frac{1}{4} \cdot 9 \right) \cdot (2)^c$$

$$1 = 2^c \quad c = 0$$

Rate law: $R = k[A]^2[B]^1[C]^0$

Behavior of polyprotic acids



Phosphoric acid is triprotic

(dissociates to produce 3 protons)

$$K_{a1} = 7.2 \times 10^{-3}$$

$$K_{a2} = 6.3 \times 10^{-8}$$

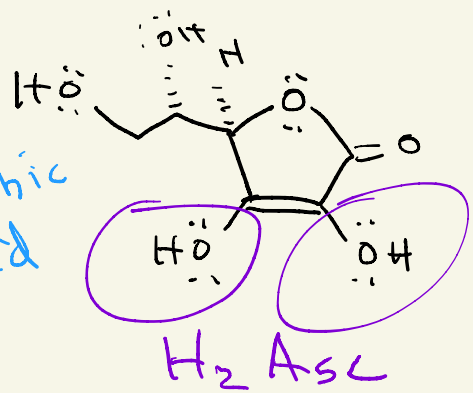
$$K_{a3} = 4.2 \times 10^{-13}$$

↓ decreasing strength

Each time an acid dissociates, the next dissociation is more difficult because it would mean removing a proton (a positive charge) from an ion.

Sample problem 18.10

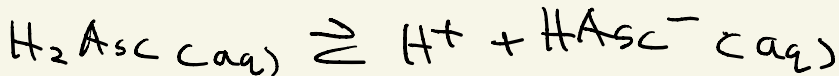
ascorbic acid



$$K_{a1} = 1.0 \times 10^{-5}$$

$$K_{a2} = 5 \times 10^{-12}$$

What is $[\text{HAsc}^-]$, $[\text{Asc}^{2-}]$ and pH? 0.050M



$$K_{a_1} = \frac{[\text{H}^+][\text{HAsc}^-]}{[\text{H}_2\text{Asc}]}$$

$K_{a_1} \gg K_{a_2} \rightarrow$ The first dissociation, being so much stronger than the second, will dictate both the concentration of its conjugate ($[\text{HAsc}^-]$) and $[\text{H}^+]$ (and therefore pH).

This problem is solved by only paying attention to the first dissociation and using those results as the starting point of the second dissociation,

$$1.0 \times 10^{-5} = \frac{(x)(x)}{(0.05 - x)}$$

assume
 $x \ll 0.05$

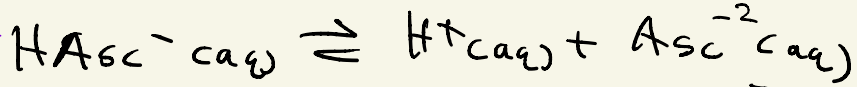
$$1.0 \times 10^{-5} = x^2 / 0.05$$

$$x^2 = 5.0 \times 10^{-7}$$

$$x = 7.1 \times 10^{-4} \quad (\text{much smaller than } 0.05)$$

$$x = [\text{H}^+] \quad \text{pH} = -\log_{10}(7.1 \times 10^{-4}) = 3.15$$

$$x = [\text{HAsc}^-] = 7.1 \times 10^{-4}$$



$$K_{a2} = 5 \times 10^{-12} = \frac{[H^{+}][A_{sc}^{-2}]}{[HA_{sc}^{-}]}$$

only look at $[H^{+}]$ produced from HA_{sc}^{-}

	HA_{sc}^{-}	H^{+}	A_{sc}^{-2}
I	7.1×10^{-4}	0	0
C	$-x$	$+x$	$+x$
E	$7.1 \times 10^{-4} - x$	x	x

$$x \lll 7.1 \times 10^{-4}$$

$$[HA_{sc}^{-}] = 7.1 \times 10^{-4}$$

$$5 \times 10^{-12} = \frac{x^2}{7.1 \times 10^{-4}}$$

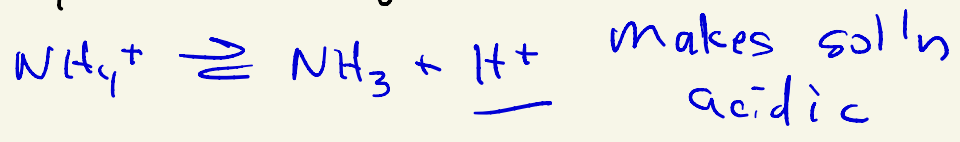
$$x = [A_{sc}^{-2}] = 6.0 \times 10^{-8}$$

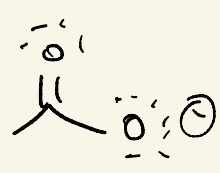
→ hold on until friday

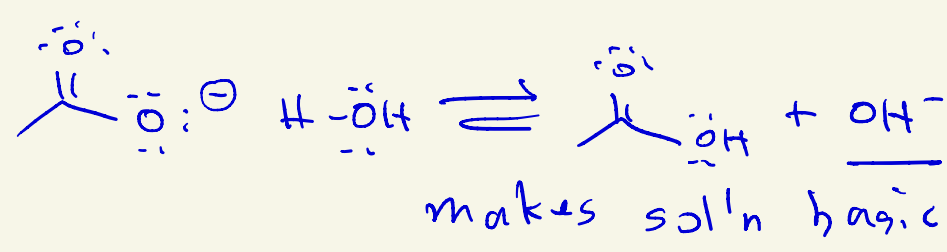
$\text{NH}_4\text{C}_2\text{H}_3\text{O}_2$ (ammonium acetate) (6)

Will this salt produce an acidic, basic, or neutral solution? $(K_{\text{a cation}} = K_{\text{b anion}})$ neutral

$\text{NH}_4^+ \rightarrow$ conjugate of a weak base



 \rightarrow conjugate of a weak acid



The answer depends on which is stronger; the acidic cation or the basic anion)

$K_{\text{a cation}} > K_{\text{b anion}} \rightarrow$ salt is acidic

$K_{\text{a cation}} < K_{\text{b anion}} \rightarrow$ salt is basic

$K_{\text{a cation}} = K_{\text{b anion}} \rightarrow$ salt is neutral