

K_a Calculations

K_a Calculations

Ionization Constant (K_a)

K_a – ionization constant for a weak acid



What is the formula for K_a ?

$$K_a = \frac{[\text{H}^+_{(aq)}][\text{C}_2\text{H}_3\text{O}_2^-_{(aq)}]}{[\text{HC}_2\text{H}_3\text{O}_{2(aq)}]}$$

K_a Calculations

Ionization Constant (K_a)

How do you suppose the K_a values of strong acids compare with weak acids?

1.Strong

- K_a values are very large
- all available H^+ ions are released into solution

2.Weak

- K_a values are small (< 1.00)
- some H^+ ions are released into solution

K_a Calculations

pK_a

How are K_a and pK_a related?

$$pK_a = -\log K_a$$

small pK_a = ionization; acid

large pK_a = ionization; acid

K_a Calculations

Example #1

- a) Calculate pK_a for acetic acid given $K_a = 1.8 \times 10^{-5}$

$$\begin{aligned}pK_a &= -\log K_a \\pK_a &= -\log (1.8 \times 10^{-5}) \\pK_a &= 4.7\end{aligned}$$

- b) Calculate K_a for ammonium ion given $pK_a = 9.24$.

$$\begin{aligned}pK_a &= -\log K_a \\K_a &= 10^{-pK_a} \\K_a &= 10^{-(9.24)} \\K_a &= 5.75 \times 10^{-10}\end{aligned}$$

K_a Calculations

Example #2

Hypoiodous acid has a pK_a of 10.6. The pK_a of hypobromous acid is 8.64. What is the chemical formula for each substance?
Which is the weaker acid?

Hypoiodous acid = HIO_(aq)
pK_a = 10.6

Hypobromous acid = HBrO_(aq)
pK_a = 8.64

Smaller pK_a = stronger acid

Check out the k_a...

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

$$K_a = 10^{-(10.6)}$$

$$K_a = 2.51 \times 10^{-11}$$

Smaller value

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

$$K_a = 10^{-(8.64)}$$

$$K_a = 2.00 \times 10^{-9}$$

K_a Calculations

Calculations using pH

Two types of calculations:

1. Calculate K_a and pK_a from the pH of its solution given initial concentration.
2. Calculate pH or $[H^+]$ of a solution given the initial concentration and K_a or pK_a .

K_a Calculations

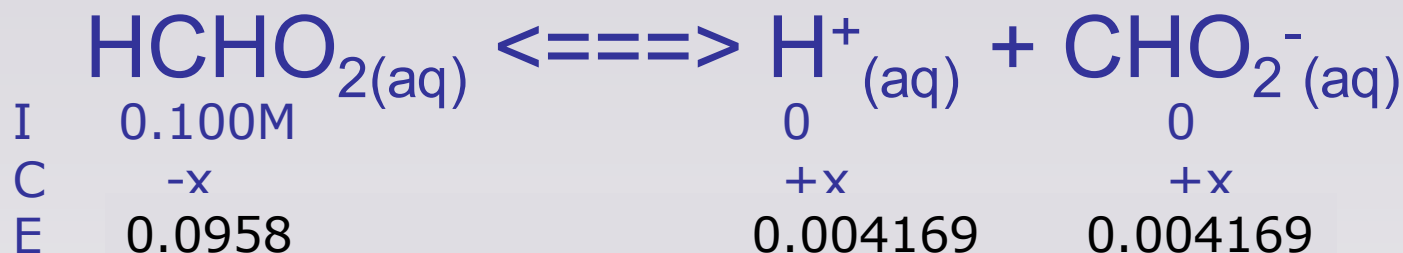
Example #3

Formic acid, HCHO_2 is a monoprotic acid. In a 0.100 M solution of formic acid, the pH is 2.38 at 25°C.

Calculate the K_a and $\text{p}K_a$ for formic acid at this temperature.

K_a Calculations

EXAMPLE #3: Formic acid, HCHO₂ is a monoprotic acid. In a 0.100 M solution of formic acid, the pH is 2.38 at 25°C. Calculate the K_a and pK_a for formic acid at this temperature.



$$[\text{H}^+] = 10^{-\text{pH}}$$

$$[\text{H}^+] = 10^{-(2.38)}$$

$$[\text{H}^+] = 0.004168694\text{M}$$

→ at equilibrium, $x = [\text{H}^+] = [\text{CHO}_2^-] = 0.004168694\text{M}$

→ $0.100 - x = 0.100 - (0.004168694) = 0.0958343$

$$K_a = \frac{[\text{H}^+(\text{aq})][\text{CHO}_2^-(\text{aq})]}{[\text{HCHO}_2(\text{aq})]}$$

$$K_a = \frac{[0.004169][0.004169]}{[0.0958]}$$

$$K_a = 0.000181425$$

$$\text{p}K_a = -\log K_a$$

$$\text{p}K_a = -\log (0.000181425)$$

$$\text{p}K_a = 3.74$$

$$\therefore k_a = 1.81 \times 10^{-4} \text{ and } \text{p}K_a = 3.74$$

K_a Calculations

Example #4

A 0.100 M solution of the weak acid HF was found to have an $[H_3O^+] = 0.0080$ M at equilibrium. Calculate the K_a and pK_a for HF.

K_a Calculations

Example #4: A 0.100 M solution of the weak acid HF was found to have an [H₃O⁺] = 0.0080 M at equilibrium. Calculate the K_a and pK_a for HF.



I	0.100M
C	-x
E	0.092

0
+x
0.008

0
+x
0.008

$$K_a = \frac{[\text{H}^+_{(\text{aq})}][\text{F}^-_{(\text{aq})}]}{[\text{HF}_{(\text{aq})}]}$$

$$\text{p}K_a = -\log K_a$$

$$K_a = \frac{[0.008][0.008]}{[0.092]}$$

$$\text{p}K_a = 3.1576$$

$$K_a = 0.000695652$$

∴ HF has a k_a of 7.0x10⁻⁴ and a pK_a of 3.2

K_a Calculations

Example #5

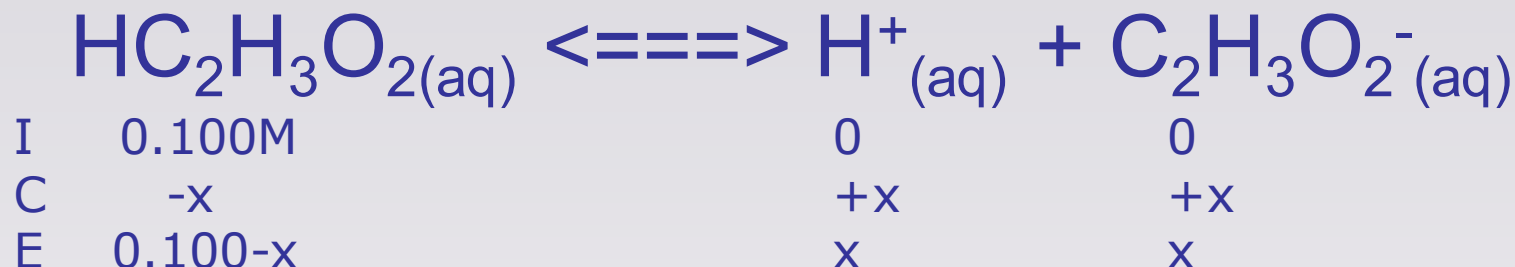
A student planned an experiment that would use 0.100 M acetic acid.

Calculate the values of $[H^+]$ and pH.

$$K_a = 1.8 \times 10^{-5}$$

K_a Calculations

Example #5: A student planned an experiment that would use 0.100 M acetic acid. Calculate the values of [H⁺] and pH. K_a = 1.8 x 10⁻⁵



$$K_a = \frac{[\text{H}^+(\text{aq})][\text{C}_2\text{H}_3\text{O}_2^-(\text{aq})]}{[\text{HC}_2\text{H}_3\text{O}_2(\text{aq})]}$$

$$1.8 \times 10^{-5} = \frac{[x][x]}{[0.100-x]}$$

Assumption $\rightarrow 0.100-x = 0.100$

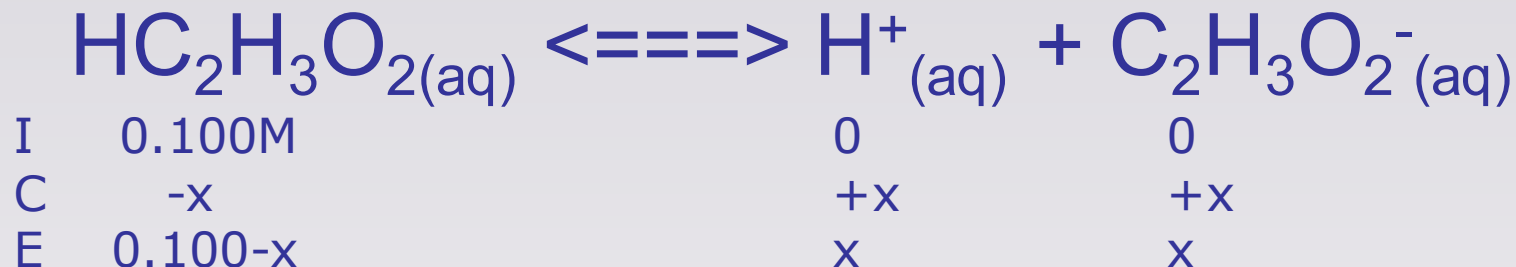
$$1.8 \times 10^{-5} (0.100) = x^2$$

$$\sqrt{1.8 \times 10^{-6}} = \sqrt{x^2}$$

$$0.00134 = x$$

K_a Calculations

Example #5 continued: A student planned an experiment that would use 0.100 M acetic acid.
Calculate the values of [H⁺] and pH. K_a = 1.8 x 10⁻⁵



$$x = 0.00134 = [\text{H}^+]$$

$$\text{pH} = -\log [\text{H}^+]$$

$$\text{pH} = -\log [0.00134]$$

$$\text{pH} = 2.87$$

$$\therefore [\text{H}^+] = 1.34 \times 10^{-3} \text{M} \text{ and } \text{pH} = 2.87$$

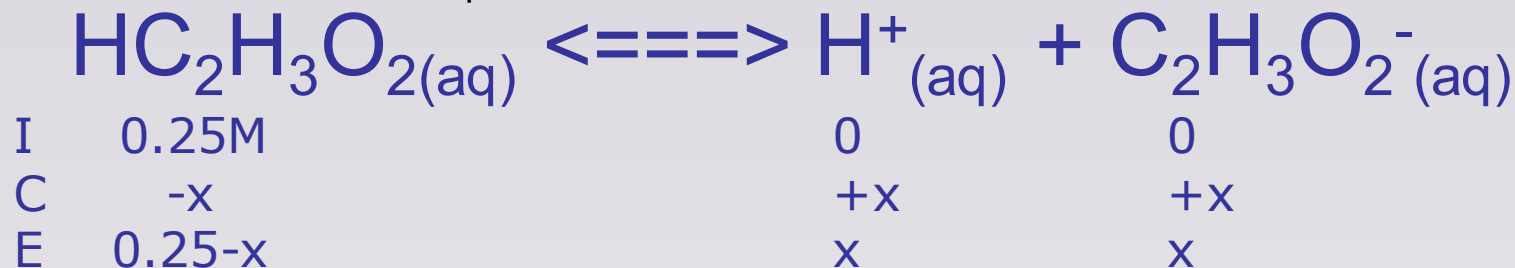
K_a Calculations

Example #6

What pH results when 0.25 mol of acetic acid is dissolved in enough water to make 1.00 L of solution? $pK_a = 4.74$

K_a Calculations

Example #6: What pH results when 0.25 mol of acetic acid is dissolved in enough water to make 1.00 L of solution? pK_a = 4.74



$$K_a = 10^{-\text{pK}_a}$$

$$K_a = 10^{-(4.74)}$$

$$K_a = 0.000018197$$

$$K_a = \frac{[\text{H}^+_{(aq)}][\text{C}_2\text{H}_3\text{O}_2^-_{(aq)}]}{[\text{HC}_2\text{H}_3\text{O}_{2(aq)}]}$$

$$1.8197 \times 10^{-5} = \frac{[x][x]}{[0.25-x]}$$

$$\text{Assumption} \rightarrow 0.25-x = 0.25$$

$$1.8197 \times 10^{-5} = \frac{[x][x]}{[0.25]}$$

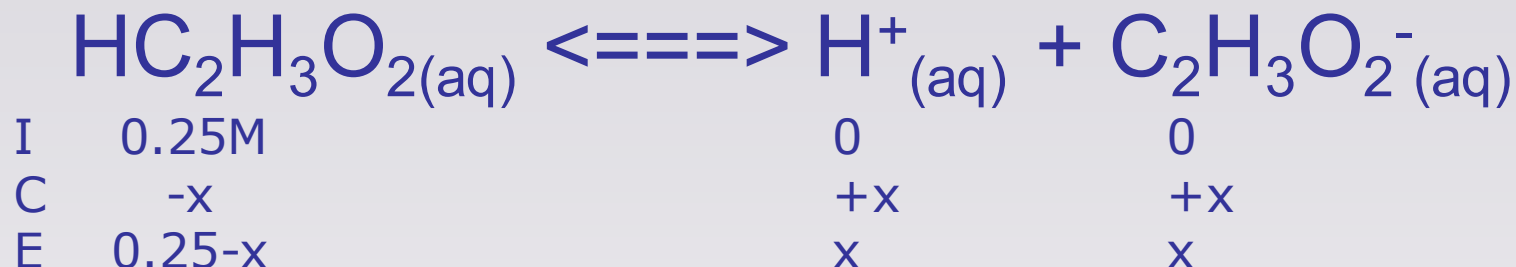
$$1.8197 \times 10^{-5} (0.25) = x^2$$

$$\sqrt{0.000004549} = \sqrt{x^2}$$

$$0.00213 = x$$

K_a Calculations

Example #6: What pH results when 0.25 mol of acetic acid is dissolved in enough water to make 1.00 L of solution? pK_a = 4.74



$$x = [\text{H}^+] = 0.00213$$

$$\text{pH} = -\log [\text{H}^+]$$

$$\text{pH} = -\log [0.00213]$$

$$\text{pH} = 2.67$$

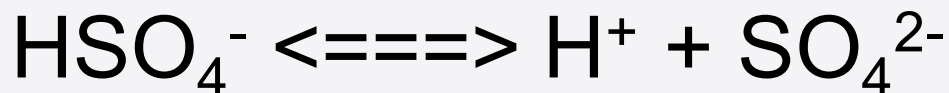
∴ the pH is 2.67

K_a Calculations

Polyprotic Acids

- acids which may release multiple H⁺ ions into solution (i.e. H₂SO₄)

Polyprotic acids release H⁺ ions into solution one step at a time.



K_a Calculations

Each H^+ released is associated with a K_a value.



Why is H_2SO_4 considered a strong acid?

Because its first proton dissociates readily in water (has a very large k_a value)

K_a Calculations

Example #7

Calculate the pH and $[\text{C}_6\text{H}_6\text{O}_6]^{2-}$ of a 0.10 M solution of $\text{H}_2\text{C}_6\text{H}_6\text{O}_{6(aq)}$.

$$K_{a1} = 7.9 \times 10^{-5}$$

$$K_{a2} = 1.6 \times 10^{-12}$$

K_a Calculations

Example #7: Calculate the pH and [C₆H₆O₆]²⁻ of a 0.10 M solution of H₂C₆H₆O_{6(aq)}.

$$K_{a1} = 7.9 \times 10^{-5}$$

$$K_{a2} = 1.6 \times 10^{-12}$$



$$\text{I} \quad 0.10\text{M}$$

$$\text{C} \quad -x$$

$$\text{E} \quad 0.09719$$

$$0$$

$$+x$$

$$0.00281$$

$$0$$

$$+x$$

$$0.00281$$

These will be used for the second H⁺ that is released

$$K_a = \frac{[\text{H}^+_{(aq)}][\text{HC}_6\text{H}_6\text{O}_6^-(aq)]}{[\text{H}_2\text{C}_6\text{H}_6\text{O}_{6(aq)}]}$$

$$7.9 \times 10^{-5} = \frac{[x][x]}{[0.10]}$$

← assumption used

$$\sqrt{7.9 \times 10^{-6}} = \sqrt{x^2}$$

$$0.002810694 = x$$

K_a Calculations

Example #7 continued: Calculate the pH and [C₆H₆O₆]²⁻ of a 0.10 M solution of H₂C₆H₆O_{6(aq)}.

$$K_{a1} = 7.9 \times 10^{-5}$$

$$K_{a2} = 1.6 \times 10^{-12}$$



I	0.00281M	0.00281	0
C	-x	+x	+x
E	0.00281-x	0.00281+x	x

$$K_a = \frac{[\text{H}^+ (\text{aq})][\text{C}_6\text{H}_6\text{O}_6^{2-} (\text{aq})]}{[\text{HC}_6\text{H}_6\text{O}_6^- (\text{aq})]}$$

$$1.6 \times 10^{-12} = \frac{[0.00281+x][x]}{[0.00281-x]}$$

$$1.6 \times 10^{-12} (0.00281-x) = 0.00281x + x^2$$

$$4.496 \times 10^{-15} - 1.6 \times 10^{-12}x = 0.00281x + x^2$$

$$0 = x^2 + 0.00281x - 4.496 \times 10^{-15}$$

$$x = 1.6 \times 10^{-12} \text{ or } -0.00281$$

K_a Calculations

Example #7 continued: Calculate the pH and [C₆H₆O₆]²⁻ of a 0.10 M solution of H₂C₆H₆O_{6(aq)}.

$$K_{a1} = 7.9 \times 10^{-5}$$

$$K_{a2} = 1.6 \times 10^{-12}$$



$$\text{I} \quad 0.00281\text{M}$$

$$0.00281$$

$$0$$

$$\text{C} \quad -x$$

$$+x$$

$$+x$$

$$\text{E} \quad 0.00281$$

$$0.00281$$

$$1.6 \times 10^{-12}$$

$$x = 1.6 \times 10^{-12}$$

$$\text{pH} = -\log [\text{H}^+]$$

$$\text{pH} = -\log [0.00281]$$

$$\text{pH} = 2.55$$

$$\therefore \text{pH} = 2.6 \text{ and } [\text{C}_6\text{H}_6\text{O}_6^{2-}(\text{aq})] = 1.6 \times 10^{-12}\text{M}$$