

A 0.10 M solution of formic acid (HCHO₂) has a pH of 2.38 at 25°C

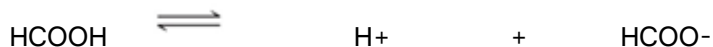
(a) Calculate *K_a* for formic acid at this temperature.

$$\text{pH} = -\log [\text{H}^+] = 2.38$$

$$\log [\text{H}^+] = -2.38$$

$$[\text{H}^+] = 10^{-2.38} = 4.1686 \text{ E-}3$$

$$K_a = \frac{[\text{H}^+][\text{HCOO}^-]}{[\text{HCOOH}]}$$



Initial	.1 M	0	0
C	- 4.1686 E-3 M	+ 4.1686 E-3 M	+ 4.1686 E-3 M
E	(0.1- 4.1686 E-3) M	4.1686 E-3 M	4.1686 E-3 M

Taking the dissociation as negligible-

$$K_a = \frac{[\text{HCOOH}]}{[\text{H}^+][\text{HCOO}^-]} \quad K_a = \frac{(4.1686 \text{ E-}3)(4.1686 \text{ E-}3)}{0.10} = 1.8 \text{ E-}4$$

(b) What percentage of the acid is ionized in this 0.10 M solution?

$$\% \text{ dissociation} = (\text{Concentration of H}^+ / \text{concentration of acid}) \times 100$$

$$= (4.1686 \text{ E-}3 / 0.10) \times 100 = 4.16\%$$

A 0.10 M solution of formic acid (HCHO₂) has a pH of 2.50 at 25°C

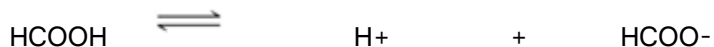
(a) Calculate *K_a* for formic acid at this temperature.

$$\text{pH} = -\log [\text{H}^+] = 2.50$$

$$\log [\text{H}^+] = -2.50$$

$$[\text{H}^+] = 10^{-2.50} = 3.16227 \text{ E-}3$$

$$K_a = \frac{[\text{H}^+][\text{HCOO}^-]}{[\text{HCOOH}]}$$



I	.1 M	0	0
C	- 3.16227 E-3 M	+ 3.16227 E-3 M	+ 3.16227 E-3 M
E	(0.1 - 3.16227 E-3) M	3.16227 E-3 M	3.16227 E-3 M

Taking the dissociation as negligible-

$$K_a = \frac{[\text{HCOOH}]}{[\text{H}^+][\text{HCOO}^-]} \quad K_a = \frac{(3.16227 \text{ E-}3)(3.16227 \text{ E-}3)}{0.10} = 1.0 \times 10^{-4}$$

(b) What percentage of the acid is ionized in this 0.10 M solution?

$$\% \text{ dissociation} = (\text{Concentration of H}^+ / \text{concentration of acid}) \times 100$$

$$= (3.16227 \text{ E-}3 / 0.10) \times 100 = 3.16 \%$$

A 0.10 M solution of formic acid (HCHO₂) has a pH of 3.30 at 25°C

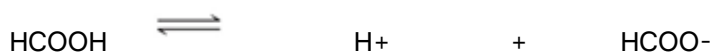
a) Calculate *K_a* for formic acid at this temperature.

$$\text{pH} = -\log [\text{H}^+] = 3.30$$

$$\log [\text{H}^+] = -3.30$$

$$[\text{H}^+] = 10^{-3.30} = 5.01187 \text{ E-}4$$

$$K_a = \frac{[\text{H}^+][\text{HCOO}^-]}{[\text{HCOOH}]}$$



I	.1 M	0	0
C	- 5.01187 E-4 M	+ 5.01187 E-4 M	+ 5.01187 E-4 M
E	(0.1 - 5.01187 E-4) M	5.01187 E-4 M	5.01187 E-4 M

Taking the dissociation as negligible-

$$K_a = \frac{[\text{HCOOH}]}{[\text{H}^+][\text{HCOO}^-]} \quad K_a = \frac{(5.01187 \text{ E-}4)(5.01187 \text{ E-}4)}{0.10} = 2.5 \text{ E-}6$$

(b) What percentage of the acid is ionized in this 0.10 M solution?

$$\% \text{ dissociation} = (\text{Concentration of H}^+ / \text{concentration of acid}) \times 100$$

$$= (5.01187 \text{ E-}4 / 0.10) \times 100 = 0.50 \%$$

A 0.10 M solution of formic acid (HCHO₂) has a pH of 3.299999999999998 at 25°C

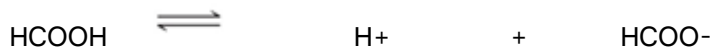
(a) Calculate *K_a* for formic acid at this temperature.

$$\text{pH} = -\log [\text{H}^+] = 3.299999999999998$$

$$\log [\text{H}^+] = -3.299999999999998$$

$$[\text{H}^+] 10^{-3.299999999999998} = 5.0118723362727209 \times 10^{-4}$$

$$K_a = \frac{[\text{H}^+][\text{HCOO}^-]}{[\text{HCOOH}]}$$



	.1 M	0	0
	- 5.0118723362727209E-4 M	+ 5.0118723362727209E-4 M	+ 5.0118723362727209E-4 M
	(0.1 - 5.0118723362727209E-4) M	5.0118723362727209E-4 M	5.0118723362727209E-4 M

Taking the dissociation as negligible-

$$K_a = \frac{[\text{HCOOH}]}{[\text{H}^+][\text{HCOO}^-]} \quad K_a = \frac{(0.1 - 5.0118723362727209 \times 10^{-4})(5.0118723362727209 \times 10^{-4})}{0.10} = 2.511886431509578 \times 10^{-6}$$

(b) What percentage of the acid is ionized in this 0.10 M solution?

$$\% \text{ dissociation} = (\text{Concentration of H}^+ / \text{concentration of acid}) \times 100$$

$$= (5.0118723362727209E-4 / 0.10000000000000001) \times 100 =$$