A 0.10 *M* solution of formic acid ($HCHO_2$) has a pH of 2.38 at 25°C

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

pH = - log [H+] = 2.38

log [H+] =-2.38

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

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

$$HCOOH \rightleftharpoons H^+ + HCOO^-$$

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-

$$K^{a} = \frac{[HCOOH]}{[H_+][HCOO_-]}$$

$$Ka = \frac{(4.1686\ E-3)(4.1686\ E-3)}{0.10} = 1.8\ E\text{-}4$$

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

% dissociation = (Concentration of H+/concentration of acid) x 100

= (4.1686 E-3/0.10) x 100 = 4. 16%

A 0.10 *M* solution of formic acid ($HCHO_2$) has a pH of 2.50 at 25°C

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

pH = - log [H+] = 2.50

log [H+] =-2.50

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

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

$$HCOOH \rightleftharpoons H^+ + HCOO^-$$

I	.1 M	0	0
C	- 3.16227 E-3 M	+ 3.16227 E-3 M	+ 3.16227 E-3 M
E	(0.2- 3.16227 E-3) M	3.16227 E-3 M	3.16227 E-3 M

Taking the dissociation as negligible-

$$K^a = \frac{[HCOOH]}{[H_+][HCOO\_]}$$

$$Ka = \frac{(3.16227\ E{-}3)(3.16227\ E{-}3)}{0.10} = 1.0 \times 10^{-4}$$

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

% dissociation = (Concentration of H+/concentration of acid) x 100

= (3.16227 E-3/0.10) x 100 = 3.16 %

A 0.10 *M* solution of formic acid ($HCHO_2$) has a pH of 3.30 at 25°C

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

pH = - log [H+] = 3.30

log [H+] =-3.30

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

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

$$HCOOH \rightleftharpoons H^+ + HCOO^-$$

I	.1 M	0	0
C	- 5.01187 E-4 M	+ 5.01187 E-4 M	+ 5.01187 E-4 M
E	(0.3- 5.01187 E-4) M	5.01187 E-4 M	5.01187 E-4 M

Taking the dissociation as negligible-

$$K^a = \frac{[HCOOH]}{[H_+][HCOO\_]}$$

$$Ka = \frac{(5.01187\ E{-}4)(5.01187\ E{-}4)}{0.10} = 2.5\ E{-}6$$

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

% dissociation = (Concentration of H+/concentration of acid) x 100

= (5.01187 E-4/0.10) x 100 = 0.50 %

A 0.10 *M* solution of formic acid ($HCHO_2$) has a pH of 3.2999999999999998at 25°C

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

pH = - log [H+] = 3.2999999999999998

log [H+] =-3.2999999999999998

[H+] 10 -3.2999999999999998 = 5.0118723362727209E-4

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

HCOOH $\rightleftharpoons$ H+ + HCOO-

	.1 M	0	0
	- 5.0118723362727209E-4 M	+ 5.0118723362727209E-4 M	+ 5.0118723362727209E-4 M
	(0.4- 5.0118723362727209E-4) M	5.0118723362727209E-4 M	5.0118723362727209E-4 M

Taking the dissociation as negligible-

$$K^{a} = \frac{[HCOOH]}{[H_+][HCOO_-]}$$

$$Ka = \frac{(5.0118723362727209E-4)(5.0118723362727209E-4)}{0.10} =$$

2.511886431509578E-6

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

% dissociation = (Concentration of H+/concentration of acid) x 100

= (5.0118723362727209E-4/0.10000000000000001) x 100 =