Percent Ionization + K_b

Percent Ionization

Percentage ionization may be determined if the K_a value of a weak acid is known.

% ionization = x 100

Percent Ionization

Example #1

What is the percentage ionization of acetic acid ($HC_2H_3O_2$) with a concentration of 0.20 M.

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

$$HC_2H_3O_{2(aq)} <==> H^+_{(aq)} + C_2H_3O_2^-{}_{(aq)}$$

	$HC_2H_3O_2$	H^+	$C_2H_3O_2^-$
I	0.20M	0	0
C	-x	+x	+x
E	0.20-x	x	x

$$K_a = 1.8 \times 10^{-5} = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

$$1.8 \times 10^{-5} = \frac{[x][x]}{[0.20-x]}$$ ←use assumption

$$1.8 \times 10^{-5} = \frac{[x][x]}{[0.20]}$$

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$$HC_2H_3O_{2(aq)} <==> H^+_{(aq)} + C_2H_3O_2^-{}_{(aq)}$$

$$1.8 \times 10^{-5} = \frac{[x][x]}{[0.20]}$$

$$1.897 \times 10^{-3} = x$$

$$\% \text{ ionization} = \frac{\text{amount of acid ionized}}{\text{amount of initial acid}} \times 100\%$$

$$\% \text{ ionization} = \frac{1.897\text{X}10^{-3}}{0.20} \times 100\%$$

$$\% \text{ ionization} = 0.94868\%$$

.: the % ionization is 0.95%

Percent Ionization

Example #1

What is the percentage ionization of acetic acid in solutions with concentration of:

i) 0.010 M 4.2%

ii) 0.0010 M 13%

Percent Ionization

Example #1

What do your calculated values indicate?

Percentage ionization increases as the solution of the weak acid becomes more dilute.

Percent Ionization

Example #2

The pH of a 0.10 M methanoic acid (HCO_2H) solution is 2.38. Calculate the percent ionization of methanoic acid.

$$HCO_2H_{(aq)} <==> H^+_{(aq)} + CO_2H^-_{(aq)}$$

	HCO_2H	H^+	CO_2H^-
I	0.10M	0	0
C	-x	+x	+x
E	0.10-x	4.16869×10^{-3}	x

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

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

$[H^+] = 4.16869\times10^{-3}$

$$\%\ \text{ionization} = \frac{\text{acid ionized}}{\text{initial acid}} \times 100\%$$

$$\%\ \text{ionization} = \frac{4.16869\times10^{-3}}{0.10} \times 100\%$$

$$\%\ \text{ionization} = 4.2\%$$

.: the % ionization is 4.2%

Percent Ionization

Example #3

Calculate K_a of acetic acid if a 0.100 M solution has a percent ionization of 1.3%.

$$HC_2H_3O_{2(aq)} \Longleftrightarrow H^+_{(aq)} + C_2H_3O_2^-{}_{(aq)}$$

	$HC_2H_3O_2$	H^+	$C_2H_3O_2^-$
I	0.100M	0	0
C	-x	+x	+x
E	0.0987	0.0013	0.0013

$$\% \text{ ionization} = \frac{\text{acid ionized}}{\text{initial acid}} \times 100\%$$

$$1.3 = \frac{\text{acid ionized}}{0.100} \times 100\%$$

$$0.0013 = \text{acid ionized}$$

$$K_a = \frac{[0.0013][0.0013]}{[0.0987]}$$

$$K_a = 1.71\text{x}10^{-5}$$

.: K_a is $1.7\text{x}10^{-5}$

K_b

K_b

K_b – dissociation constant for weak bases

$$NH_3 + H_2O \Longleftrightarrow NH_4^+ + OH^-$$

The K_b, the stronger the base.

The K_b, the weaker the base.

K_b

Similar to K_a values, pK_b values can be calculated.

The larger the pK_b value, the weaker the base and the smaller the pK_b, the stronger the base.

K_b

How are K_a and K_b related?

Write out the equilibrium eq^n for acetic acid.

$$\cancel{HC_2H_3O_{2(aq)}} + H_2O_{(l)} <==> H_3O^+_{(aq)} + \cancel{C_2H_3O_2^-{}_{(aq)}}$$

Write out the equilibrium eq^n for the conjugate base of acetic acid.

$$\cancel{C_2H_3O_2^-{}_{(aq)}} + H_2O_{(l)} <==> \cancel{HC_2H_3O_{2(aq)}} + OH^-_{(aq)}$$

Perform Hess' Law.

$$2\ H_2O_{(l)} <==> H_3O^+_{(aq)} + OH^-_{(aq)}$$

$$\boxed{K_a \times K_b = K_w}$$

K_b

RECALL: Equilibrium Law

What is the equilibrium law of the sum of the following reactions?

$$2\,N_{2(g)} + O_{2(g)} \Longleftrightarrow \cancel{2\,N_2O_{(g)}} \qquad k_{eq} = \frac{[N_2O_{(g)}]^2}{[N_{2(g)}]^2[O_{2(g)}]}$$

$$\cancel{2\,N_2O_{(g)}} + 3\,O_{2(g)} \Longleftrightarrow 4\,NO_{2(g)} \qquad k_{eq} = \frac{[NO_{2(g)}]^4}{[N_2O_{(g)}]^2[O_{2(g)}]^3}$$

$$k_{eq} = \frac{[NO_{2(g)}]^4}{[N_{2(g)}]^2[O_{2(g)}]^4}$$

$$K_{eq\ final} = \frac{\cancel{[N_2O_{(g)}]^2}}{[N_{2(g)}]^2[O_{2(g)}]} \times \frac{[NO_{2(g)}]^4}{\cancel{[N_2O_{(g)}]^2}[O_{2(g)}]^3}$$

$$= \frac{[NO_{2(g)}]^4}{[N_{2(g)}]^2[O_{2(g)}]^4}$$

K_b

RECALL: Equilibrium Law

When chemical equilibria are added together, the equilibrium constants are multiplied together.

$$K_{eq\ final\ rxn} = K_{eq\ rxn\ 1} \times K_{eq\ rxn\ 2}$$

K_b Calculations

Two types of calculations may also be completed:

1) Calculate the values of K_b and pK_b from the pH of a solution of a weak base of known initial concentration.

2) Calculate the pH of a solution where pK_b and initial concentration are known.

K_b

Example #4

Methylamine, CH_3NH_2, is one of several substances that give herring brine its pungent odor. In 0.100 M CH_3NH_2, the pH is 11.80. What is the K_b of methylamine?

$$CH_3NH_{2(aq)} + H_2O_{(l)} <==> OH^-_{(aq)} + CH_3NH_3{}^+_{(aq)}$$

	CH_3NH_2	OH^-	$CH_3NH_3^+$
I	0.100M	0	0
C	-x	+x	+x
E	0.0937	0.0063	0.0063

pH + pOH = 14

pOH = 14 – 11.80

pOH = 2.2

$[OH^-] = 10^{-pOH}$

$[OH^-] = 10^{-(2.2)}$

$[OH^-] = 6.30957\text{x}10^{-3}$

$$K_b = \frac{[0.0063][0.0063]}{[0.0937]}$$

$K_b = 4.235859\text{x}10^{-4}$

.: K_b is $4.24\text{x}10^{-4}$

K_b

Example #5

$C_{17}H_{19}NO_3$ Morphine is an alkaloid (an alkaline compound obtained from plants), which is a weak base. The pH of 0.010 M morphine is 10.10. Calculate K_b and pK_b morphine.

$$\text{morphine}_{(aq)} + H_2O_{(l)} \rightleftharpoons OH^-_{(aq)} + \text{morphine-H}^+_{(aq)}$$

	morphine		OH^-	morphine-H^+
I	0.010M		0	0
C	-x		+x	+x
E	9.8741×10^{-3}		1.2589×10^{-4}	1.2589×10^{-4}

$pOH = 14 - 10.10$

$pOH = 3.9$

$[OH^-] = 10^{-pOH}$

$[OH^-] = 10^{-(3.9)}$

$[OH^-] = 1.2589\times10^{-4}$

$$K_b = \frac{[1.2589\times10^{-4}][1.2589\times10^{-4}]}{[9.8741\times10^{-3}]} \qquad K_b = 1.605\times10^{-6}$$

K_b

Example #5

$C_{17}H_{19}NO_3$ Morphine is an alkaloid (an alkaline compound obtained from plants), which is a weak base. The pH of 0.010 M morphine is 10.10. Calculate K_b and pK_b morphine.

$$\text{morphine}_{(aq)} + H_2O_{(l)} <==> OH^-_{(aq)} + \text{morphine-H}^+_{(aq)}$$

	morphine	OH⁻	morphine-H⁺
I	0.010M	0	0
C	-x	+x	+x
E	9.8741×10^{-3}	1.2589×10^{-4}	1.2589×10^{-4}

$K_b = 1.605 \times 10^{-6}$

$pK_b = -\log K_b$

$pK_b = -\log (1.605 \times 10^{-6})$

$pK_b = 5.79$

.: $pK_b = 5.8$, $k_b = 1.6 \times 10^{-6}$

K_b

Example #6

Calculate the values of pH, pOH and $[OH^-]$ of a 0.20 M solution of ammonia. K_b of ammonia is 1.8×10^{-5}

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons OH^-_{(aq)} + NH_4^+{}_{(aq)}$$

	$NH_{3(aq)}$	$OH^-_{(aq)}$	$NH_4^+{}_{(aq)}$
I	0.20M	0	0
C	-x	+x	+x
E	0.20-x	x	x

$$K_b = \frac{[OH^-_{(aq)}][NH_4^+{}_{(aq)}]}{[NH_{3(aq)}]}$$

$$1.8 \times 10^{-5} = \frac{[x][x]}{[0.20-x]}$$

$$1.8 \times 10^{-5} = \frac{[x][x]}{[0.20]}$$ ←use assumption

$$1.897 \times 10^{-3} = x$$

pOH = -log $[OH^-]$

pOH = -log $[1.897 \times 10^{-3}]$

pOH = 2.72

pH = 14-2.72

pH = 11.28

.: pH = 11.3, pOH = 2.7, $[OH^-] = 1.9 \times 10^{-3}$ M