

Percent Ionization + K_b

Percent Ionization

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

$$\% \text{ ionization} = \boxed{} \times 100$$

Percent Ionization

Example #1

What is the percentage ionization of acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$) with a concentration of 0.20 M.

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



I	0.20M	0	0
C	-x	+x	+x
E	0.20-x	x	x

$$K_a = 1.8 \times 10^{-5} = \frac{[\text{H}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]}$$

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

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

Percent Ionization

Example #1

What is the percentage ionization of acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$) with a concentration of 0.20 M.

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



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

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

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

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

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

\therefore 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.



I 0.10M

0

0

C -x

+x

+x

E 0.10-x

4.16869x10⁻³

x

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

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

$$[\text{H}^+] = 4.16869 \times 10^{-3}$$

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

$$\% \text{ ionization} = \frac{4.16869 \times 10^{-3}}{0.10} \times 100\%$$

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

∴ 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%.



I 0.100M

0

0

C -x

+x

+x

E 0.0987

0.0013

0.0013

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

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

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

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

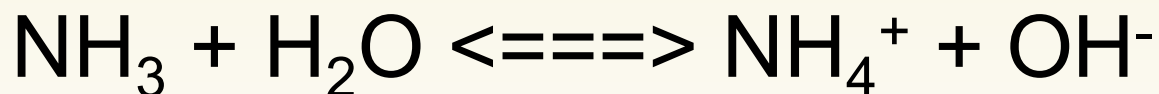
$$K_a = 1.71 \times 10^{-5}$$

$$\therefore K_a \text{ is } 1.7 \times 10^{-5}$$

K_b

K_b

K_b – dissociation constant for weak bases



The K_b , the stronger the base.

The K_b , the weaker the base.

K_b

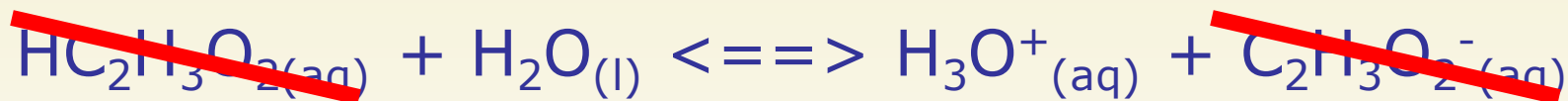
Similar to K_a values, pK_b values can be calculated.

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

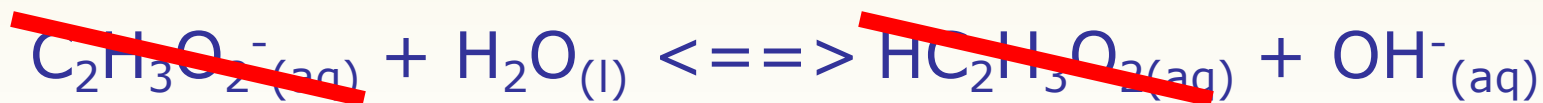
K_b

How are K_a and K_b related?

Write out the equilibrium eqⁿ for acetic acid.



Write out the equilibrium eqⁿ for the conjugate base of acetic acid.



Perform Hess' Law.

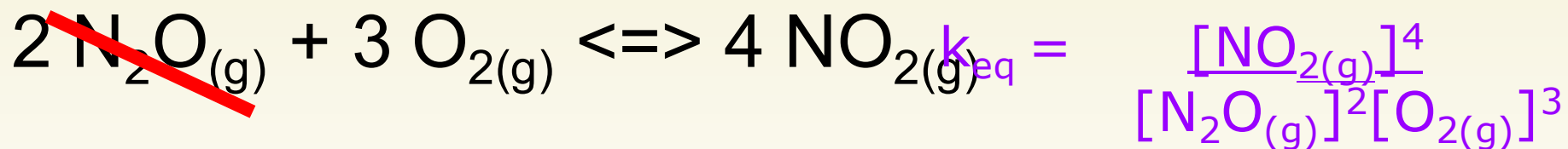
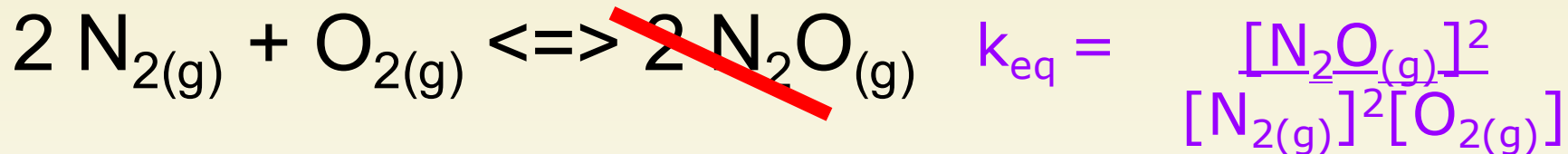


$$\boxed{K_a \times K_b = K_w}$$

K_b

RECALL: Equilibrium Law

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



$$K_{\text{eq final}} = \frac{[\cancel{\text{N}_2\text{O}_{(g)}}]^2}{[\text{N}_{2(g)}]^2[\text{O}_{2(g)}]} \times \frac{[\text{NO}_{2(g)}]^4}{[\cancel{\text{N}_2\text{O}_{(g)}}]^2[\text{O}_{2(g)}]^3}$$
$$= \frac{[\text{NO}_{2(g)}]^4}{[\text{N}_{2(g)}]^2[\text{O}_{2(g)}]^4}$$

K_b

RECALL: Equilibrium Law

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

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



K_b Calculations

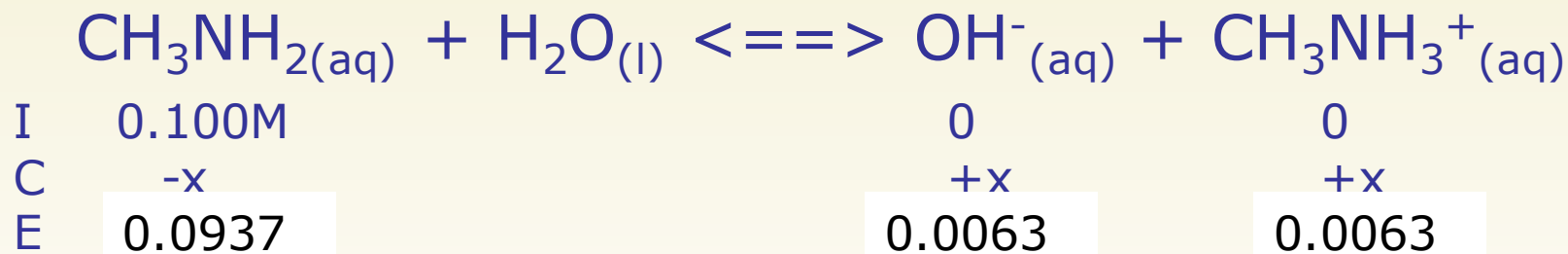
Two types of calculations may also be completed:

- 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?



$$\text{pH} + \text{pOH} = 14$$

$$\text{pOH} = 14 - 11.80$$

$$\text{pOH} = 2.2$$

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

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

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

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

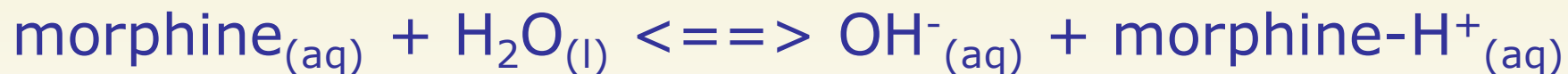
$$K_b = 4.235859 \times 10^{-4}$$

$$\therefore K_b \text{ is } 4.24 \times 10^{-4}$$

K_b

Example #5

C₁₇H₁₉NO₃ 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.



I 0.010M

C -x

E 9.8741x10⁻³

0

+x

1.2589x10⁻⁴

0

+x

1.2589x10⁻⁴

$$\text{pOH} = 14 - 10.10 \quad [\text{OH}^{-}] = 10^{-\text{pOH}}$$

$$\text{pOH} = 3.9 \quad [\text{OH}^{-}] = 10^{-(3.9)}$$

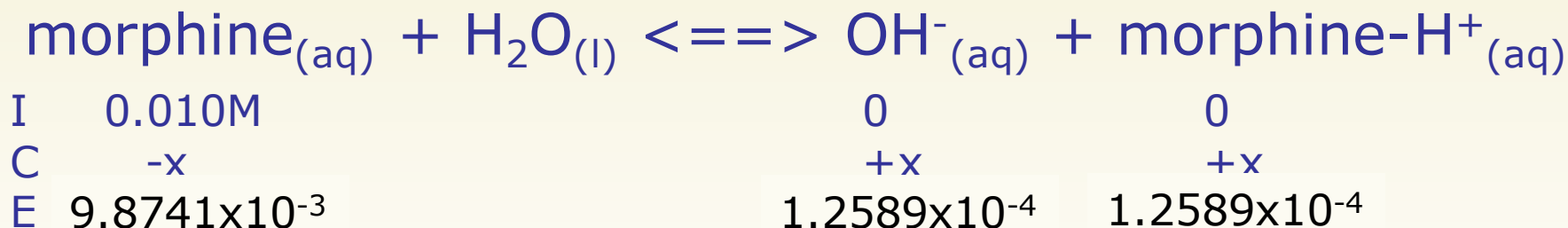
$$[\text{OH}^{-}] = 1.2589 \times 10^{-4}$$

$$K_b = \frac{[1.2589 \times 10^{-4}][1.2589 \times 10^{-4}]}{[9.8741 \times 10^{-3}]} \quad K_b = 1.605 \times 10^{-6}$$

K_b

Example #5

C₁₇H₁₉NO₃ 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.



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

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

$$\text{p}K_b = -\log (1.605 \times 10^{-6})$$

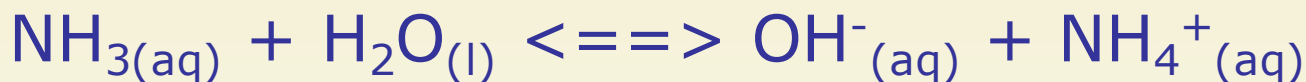
$$\text{p}K_b = 5.79$$

$$\therefore \text{p}K_b = 5.8, K_b = 1.6 \times 10^{-6}$$

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 1.8x10⁻⁵**



I 0.20M

0

0

C -x

+x

+x

E 0.20-x

x

x

$$K_b = \frac{[\text{OH}^{-}_{(\text{aq})}][\text{NH}_4^{+}_{(\text{aq})}]}{[\text{NH}_{3(\text{aq})}]}$$

$$\text{pOH} = -\log [\text{OH}^{-}]$$

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

$$\text{pOH} = 2.72$$

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

$$[0.20-x]$$

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

$$[0.20] \leftarrow \text{use assumption}$$

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

$$\text{pH} = 14 - 2.72$$

$$\text{pH} = 11.28$$

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