

Quadrant II – Transcript and Related Materials

Programme: Bachelor of Science (First Year)

Subject: Chemistry

Paper Code: CHC 102

Paper Title: Physical Chemistry and Organic Chemistry

Unit: Ionic Equilibria

Module Name: Problems on degree of dissociation, ionization constant, ionic product

Module No: 17

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DEGREE OF IONISATION OR DISSOCIATION

The fraction of the total number of molecules of the electrolyte that dissociates into ions is called the degree of ionization or dissociation.

- It is usually denoted by α
- *Degree of dissociation (α) = $\frac{\text{No. of molecules dissociated}}{\text{Total no. of molecules taken}}$*

IONISATION CONSTANT

- Ionisation and dissociation constant of the acid

$$K_a = \frac{C\alpha^2}{(1 - \alpha)}$$

- Ionisation and dissociation constant of the base

$$K_b = \frac{C\alpha^2}{(1 - \alpha)}$$

C is the initial concentration of the acid or base & α is the degree of dissociation.

IONIC PRODUCT OF WATER

- It is denoted by K_w
- $K_w = [H^+] [OH^-]$

RELATION BETWEEN K_a and K_b

For conjugate ACID-BASE PAIR,

$$K_w = K_a \times K_b$$

PROBLEMS

1. The degree of ionisation of a 0.1 M bromoacetic acid solution is 0.132.

Calculate the pK_a of bromoacetic acid.

Solution:

Given: $\alpha = 0.132$, $C = 0.1$ M

$pK_a = ???$ $K_a = ?$

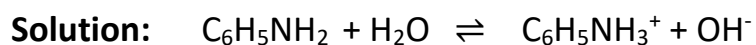
Degree of ionization is given by, $K_b = \frac{C\alpha^2}{(1 - \alpha)} = \frac{0.1 \times (0.132)^2}{(1 - 0.132)}$

$$K_b = 2 \times 10^{(-3)}$$

pK_a and K_a are related as,

$$\begin{aligned} \text{pK}_a &= -\log K_a \\ &= -\log (2 \times 10^{(-3)}) \\ \text{pK}_a &= 2.6974 \end{aligned}$$

2. The ionization constant of 0.001 M aniline is $4.27 \times 10^{(-10)}$. Calculate the degree of ionization of aniline in the solution. Also calculate the ionization constant of the conjugate acid of aniline.



Degree of ionization is given by, $K_b = \frac{C\alpha^2}{(1-\alpha)}$

Given: C = 0.001 M

$$K_b = 4.27 \times 10^{(-10)}$$

$$\alpha = ?$$

$$K_b = \frac{C\alpha^2}{(1-\alpha)}$$

$$K_b = C\alpha^2 \text{ (Assuming that degree of dissociation is very small)}$$

$$\begin{aligned} \alpha &= \sqrt{\frac{K_b}{C}} \\ &= \sqrt{\frac{4.27 \times 10^{(-10)}}{0.001}} = 6.53 \times 10^{(-4)} \end{aligned}$$

For a pair of conjugate acid and base,

$$pK_a + pK_b = 14$$

$$pK_b = -\log K_b$$

$$pK_b = -\log (4.27 \times 10^{-10})$$

we get, **$pK_b = 9.38$**

Substituting $pK_b = 9.38$ in equation $pK_a + pK_b = 14$

$$\text{We get, } pK_a + 9.38 = 14$$

$$pK_a = 4.62$$

we know that, $K_a = \text{Antilog} (-pK_a)$

$$K_a = \text{Antilog} (-4.62) = \mathbf{2.399 \times 10^{-5}}$$

3. The ionisation constant of dimethylamine is 5.4×10^{-4} . Calculate its degree of ionisation in its 0.02 M solution. What percentage of dimethylamine is ionised if the solution is also having 0.1 M in NaOH?

Solution:

Using: $K_b = \frac{C\alpha^2}{(1-\alpha)} = C\alpha^2$ (Assuming that degree of dissociation is very small)

Given: $K_b = 5.4 \times 10^{-4}$

$C = 0.02 \text{ M}$

$$\alpha = \sqrt{\frac{K_b}{c}} = \sqrt{\frac{5.4 \times 10^{-4}}{0.02}} = 0.164$$

Suppose if x is the amount of dimethyl aniline dissociated in presence of 0.1 M NaOH then,



Initially 0.02 M

After dissociation	0.02 - x	x	0.1 + x
	= 0.02		= 0.1

$$K_b = \frac{[(\text{CH}_3)_2\text{NH}^+\text{OH}^-][\text{OH}^-]}{(\text{CH}_3)_2\text{NH}} = \frac{x(0.1)}{0.02}$$

Degree of dissociation can also be expressed as, $\alpha = \frac{\text{Amount dissociated}}{\text{Initial concentration}}$

$$\alpha = \frac{x}{0.02}$$

$$K_b = \frac{x(0.1)}{0.02}$$

substituting $\alpha = \frac{x}{0.02}$ in the above equation, we get

$$K_b = \alpha \times 0.1$$

$$\alpha = \frac{K_b}{0.1} = \frac{5.4 \times 10^{-4}}{0.1}$$

$$\alpha = 5.4 \times 10^{-3}$$

% of dimethylamine ionized if the solution is also having 0.1 M in NaOH, will be therefore $5.4 \times 10^{-3} \times 100$ i.e. 0.54 %.

4. The ionization constant of HF is 6.8×10^{-4} . Calculate the ionization constant of the corresponding conjugate base.

Solution:

For conjugate ACID-BASE PAIR,

$$K_w = K_a \times K_b$$

Given: $K_a = 6.8 \times 10^{-4}$

$$K_w = 1.0 \times 10^{-14}$$

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{6.8 \times 10^{-4}} = 1.5 \times 10^{-11}$$