Physical pharmacy I
2nd stage
Ionic equilibria
2025-2026

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Outlines

- Objectives
- Theories
- Acid-base equilibria
- Calculation of pH, acidity constants
- The effect of ionic strength.

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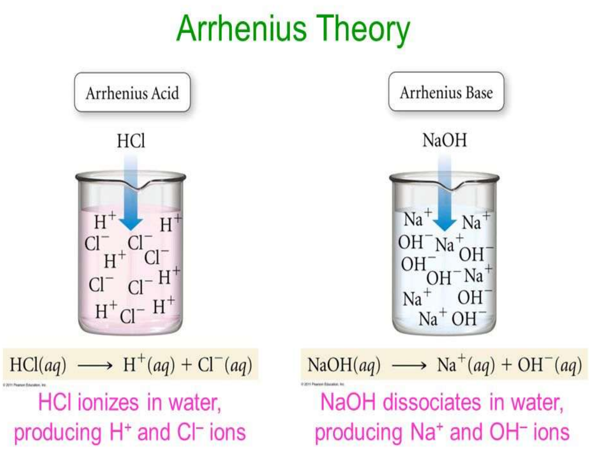
Objectives

- Define of acids and bases
- Concept of Sörensen's pH scale.
- Understanding different terminology such as Ampholytes, Aprotic , etc
- Ionization of Polyprotic electrolytes.
- pKa and pH calculation of aqueous solutions with different composition

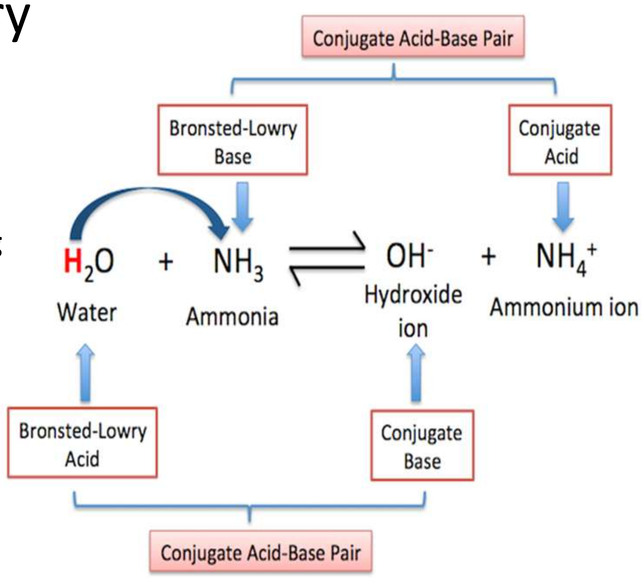
Theories

- **Arrhenius Theory**
- Arrhenius defined
  - an **acid** as a substance that liberates hydrogen ions
  - and a **base** as a substance that supplies hydroxyl ions on dissociation in aqueous media.



Brönsted–Lowry theory

- According to the Brönsted–Lowry theory,
  - an acid is a substance, charged or uncharged, that is capable of **donating** a proton,
  - and a base is a substance, charged or uncharged, that is capable of **accepting** a proton from an acid.



- The strength of an acid or a base varies with the ***solvent***.
- HCl is a **strong** acid but it is a **weak** acid in glacial acetic acid.
- Acetic acid, which is a **weak** acid, is a **strong** acid in liquid ammonia.

- Consequently, the **strength** of an acid depends:
  1. not only on its ability to give up a proton
  2. but also on the ability of the **solvent** to accept the proton from the acid.

  This is called ***the basic strength of the solvent***.

In the Brönsted–Lowry classification, **acids** and **bases** may be

- anions such as HSO_4^- and CH_3COO^-,
- cations such as NH_4^+ and H_3O^+,
- or neutral molecules such as HCl and NH_3.
- Water can act as either an acid or a base and thus is **amphiprotic**.

Solvent classification

1. **<u>A protophilic or basic solvent</u>**
   - ✓ is one that is capable of **accepting** protons from the solute. Eg. acetone, ether, and liquid ammonia.
2. **A protogenic solvent**
   - ✓ is a **proton-donating** compound and is represented by acids such as formic acid, acetic acid, sulfuric acid, liquid HCl, and liquid HF.
3. **<u>Amphiprotic solvents</u>**
   - ✓ act as both proton acceptors and proton donors, and this class includes water and alcohols.
4. **<u>Aprotic solvents,</u>**
   - ✓ such as the hydrocarbons, neither accept nor donate protons, and, being neutral in this sense, they are useful for studying the reactions of acids and bases free of solvent effects.

Proteolytic reactions or protolysis.

- Acid-base reactions occur when an acid reacts with a base to form a new acid and a new base, called conjugates. So it is the reaction that involve a transfer of a proton, and are known as **protolytic** reactions or **protolysis**. Proton transfer reactions are also known as **protonation–deprotonation reactions**.

$$\underset{\text{Acid}_1}{HCl} + \underset{\text{Base}_2}{H_2O} \rightarrow \underset{\text{Acid}_2}{H_3O^+} + \underset{\text{Base}_1}{Cl^-}$$

hydronium ion.

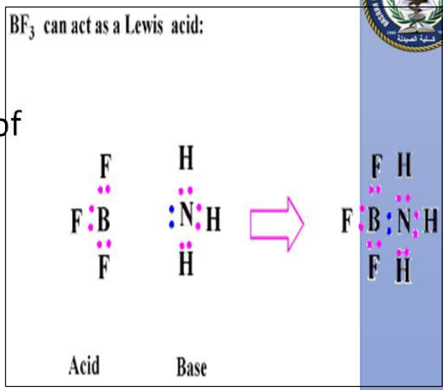
Several examples illustrate these types of reactions, as shown in Table 7-1.

TABLE 7–1. *Examples of Acid–Base Reactions*

	$Acid_1$		$Base_2$		$Acid_2$		$Base_1$
Neutralization	NH_4^+	+	OH^-	=	H_2O	+	NH_3
Neutralization	H_3O^+	+	OH^-	=	H_2O	+	H_2O
Neutralization	HCl	+	NH_3	=	NH_4^+	+	Cl^-
Hydrolysis	H_2O	+	CH_3COO^-	=	CH_3COOH	+	OH^-
Hydrolysis	NH_4^+	+	H_2O	=	H_3O^+	+	NH_3
Displacement	HCl	+	CH_3COO^-	=	CH_3COOH	+	Cl^-

Lewis Electronic Theory.

- According to the Lewis theory, an acid is a molecule or an ion that **accepts** an **electron pair** to form a covalent bond.
- A base is a substance that **provides** the pair of **unshared electrons** by which the base coordinates with an acid.
- Lewis acid eg boron trifluoride and aluminum chloride,
- Lewis base eg amines, ethers, and carboxylic acid anhydrides, are classified as bases according to the Lewis definition.



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- The Lewis acid theory is widely used for describing the mechanism of many organic and inorganic reactions.
- It is referred to simply as a form of electron sharing rather than as acid–base reactions.
- It will be important in **solubility** and **complexation**.
- *The Brønsted–Lowry nomenclature is* ***particularly useful*** *for describing ionic equilibria and is used extensively in this chapter*

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Sörensen's pH

- H ion concentration of a solution varies from approximately 1 in a 1 M solution of a strong acid to about 1 × 10-14 in a 1 M solution of a strong base
- The pH of a solution can be considered in terms of a numeric scale having values from 0 to 14, which expresses in a quantitative way the degree of acidity (7 to 0) and alkalinity (7-14).
- The value 7 at which the hydrogen and hydroxyl ion concentrations are about equal at room temperature is referred to as the neutral point, or neutrality. The neutral pH at 0°C is 7.47, and at 100°C it is 6.15.

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Ionic Equilibria

- Equilibrium can be deffined as a balance between two opposing forces or actions. TRUE or FALSE?
- This statement **does not imply** cessation of the opposing reactions. Rather, it suggests a dynamic equality between the velocities of the two reactions. Tor F?
- Equilibrium is the condition where the standard free energy difference between the two sides of reaction equation (7-4) is zero ($\Delta G° = 0$). T or F ?

Ionic Equilibria

The arrows pointing in the forward and reverse directions indicate that reactions are proceeding to the right and left simultaneously

$$\underset{\text{Acid}_1}{HAc} + \underset{\text{Base}_2}{H_2O} \rightleftharpoons \underset{\text{Acid}_2}{H_3O^+} + \underset{\text{Base}_1}{Ac^-} \quad (7\text{-}4)$$

Rate of forward $= k_1 \times [HAc]^1 \times [H_2O]^1$

Rate of backward $= k_2 \times [H_3O^+]^1 \times [Ac^-]^1$

At equilibrium $k_1 \times [HAc] \times [H_2O] = k_2 \times [H_3O^+] \times [Ac^-]$

$$k = \frac{k_1}{k_2} = \frac{[H_3O^+][Ac^-]}{[HAc][H_2O]}$$

- Mwt of H2O=18.02g/L
- In dilute solutions of acetic acid, water is in suffcient excess to be regarded as constant .
- 1 L of H2O at 25°C weighs 997.07 g,
- Thus, the conc. Of H2O= 997.07/18.02 = 55.3 M).
- [H2O]=55.3 M

$$K_a = 55.3k = \frac{[H_3O^+][Ac^-]}{[HAc]}$$

Ka : ionization constant or or the dissociation constant of acetic acid

At large conc. Of water (diluent)

$$K_a = 55.3k = \frac{[H_3O^+][Ac^-]}{[HAc]}$$

$$\begin{array}{lcccc} HAc + H_2O & \rightleftharpoons & H_3O^+ & + & Ac^- \\ (c - x) & & x & & x \end{array}$$

$$K_a = \frac{x^2}{c - x}$$

where c is large in comparison with x. The term c – x can be replaced by c without appreciable error, giving the equation

$$K_a \cong \frac{x^2}{c}$$

$$x^2 = K_a c$$

$$x = [H_3O^+] = \sqrt{K_a c}$$

EXAMPLE 7-1

- In a liter of a 0.1 M solution, acetic acid was found by conductivity analysis to dissociate into 1.32 × 10−3 moles each of hydronium and acetate ions at 25°C. What is the acidity (or dissociation) constant Ka for acetic acid?
- H.W

$HB + H_2O \rightleftharpoons B^- + H_3O$

$B^- + H_2O \rightleftharpoons BH + OH$

Relationship Between K_a and K_b

A simple relationship exists between the dissociation constant of a weak acid HB and that of its conjugate base B−, or between BH+ and B, when the solvent is **amphiprotic**. This can be obtained by multiplying the following equations:

$$K_a = \frac{[H_3O^+][B^-]}{[HB]} \quad \times \quad K_b = \frac{[OH^-][HB]}{[B^-]}$$

$$K_a K_b = \frac{[H_3O^+][B^-]}{[HB]} \cdot \frac{[OH^-][HB]}{[B^-]} = [H_3O^+][OH^-] = K_w$$

$K_a = K_w/K_b$ $\quad$ $K_b = K_w/K_a$

K_w. known as the *autoprotolysis constant,* or the *ion product* of water

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EXAMPLE 7-4
Calculate Ka

- Ammonia has Kb = 1.74 × 10−5 at 25°C. Calculate Ka for its conjugate acid, NH4+. We have

$$K_a = \frac{K_w}{K_b} = \frac{1.00 \times 10^{-14}}{1.74 \times 10^{-5}}$$
$$= 5.75 \times 10^{-10}$$

Ionization of Polyprotic Electrolytes.

- Acids that donate a **single** proton and bases that accept a single proton are called **monoprotic** electrolytes.
- A polyprotic (polybasic) acid is one that is capable of donating two or more protons, and a polyprotic base is capable of accepting two or more protons.
- A diprotic (dibasic) acid, such as carbonic acid, ionizes in **two** stages, and a triprotic (tribasic) acid, such as phosphoric acid, ionizes in **three** stages.
- In any polyprotic electrolyte, the primary protolysis is greatest, and succeeding stages become less complete at any given acid concentration.

Ionization of Polyprotic Electrolytes.

$$H_3PO_4 + H_2O = H_3O^+ + H_2PO_4^-$$

$$\frac{[H_3O^+][H_2PO_4^-]}{[H_3PO_4]} = K_1 = 7.5 \times 10^{-3}$$

$$H_2PO_4^- + H_2O = H_3O^+ + HPO_4^{2-}$$

$$\frac{[H_3O^+][HPO_4^{2-}]}{[H_2PO_4^-]} = K_2 = 6.2 \times 10^{-8}$$

$$HPO_4^{2-} + H_2O = H_3O^+ + PO_4^{3-}$$

$$\frac{[H_3O^+][PO_4^{3-}]}{[HPO_4^{2-}]} = K_3 = 2.1 \times 10^{-13}$$

$$PO_4^{3-} + H_2O \rightleftharpoons HPO_4^{2-} + OH^-$$

$$K_{b1} = \frac{[HPO_4^{2-}][OH^-]}{[PO_4^{3-}]} = 4.8 \times 10^{-2}$$

$$HPO_4^{2-} + H_2O \rightleftharpoons H_2PO_4^- + OH^-$$

$$K_{b2} = \frac{[H_2PO_4^-][OH^-]}{[HPO_4^{2-}]} = 1.6 \times 10^{-7}$$

$$H_2PO_4^- + H_2O \rightleftharpoons H_3PO_4 + OH^-$$

$$K_{b3} = \frac{[H_3PO_4][OH^-]}{[H_2PO_4^-]} = 1.3 \times 10^{-12}$$

Ampholytes

Zwitterion

- A species that can function either as an acid or as a base is called an ampholyte and is said to be amphoteric in nature. Amino acids and proteins are ampholytes of particular interest in pharmacy.
- glycine hydrochloride is dissolved in water,:

$$^{+}NH_3CH_2COOH + H_2O \rightleftharpoons H_3O^{+} + {}^{+}NH_3CH_2COO^{-}$$

Amphoteric compound

react as acid with H2O → $^{+}NH_3CH_2COO^{-} + H_2O \rightleftharpoons H_3O^{+} + NH_2CH_2COO^{-}$

react as base with H2O → $^{+}NH_3CH_2COO^{-} + H_2O \rightleftharpoons OH^{-} + {}^{+}NH_3CH_2COOH$

- The amphoteric species +NH3CH2COO– is also called a zwitterion because?
- Ans.: it carries both negative and positive charge; the whole molecule is electrically neutral

Isoelectric point (IEP)

- It is pH at which zwitterion concentration at maximum
- It has been used for determination of protein and amino acids.
- The molecules have ------------- solubility at IEP.
- The net charge of a molecule at its IEP is -----------.
- Charge of a molecule at pH above its IEP is ----------- and below is -----------.

e.g. pH of milk 6.6 (casein IEP=4.6)

Solutions containing Strong acids

- **Strong acid** concentration of H is equal to initial concentration of acid. Thus, they are considered to ionize fully in aqueous solutions.

Solutions containing Only a Weak Acid

- Conc. of base (Cb) is zero.
- [H3O+] is generally much greater than [OH−]

$$[H_3O^+] = \frac{-K_a + \sqrt{K_a^2 + 4K_aC_a}}{2}$$

In many instances ,Ca is much greater than [H3O+] , and the eq simplifies further to :

$$\boxed{[H_3O^+] = \sqrt{K_aC_a}}$$

EXAMPLE 7-12 calculate the pH

- Calculate the pH of a 0.01 M solution of salicylic acid, which has Ka = 1.06 × 10−3 at 25°C.

(a) Using equation (7-102),

$$\boxed{[H_3O^+] = \sqrt{K_aC_a}}$$

$$[H_3O^+] = \sqrt{(1.06\times10^{-3})\times(1.0\times10^{-2})}$$
$$= 3.26\times10^{-3}\ \text{M} \quad =0.00326\text{M}$$

Compare to the conc of acid, The approximation that Ca >> [H3O+] is not valid.

(b) Using equation (7-101), we find

$$[H_3O^+] = \frac{-K_a + \sqrt{K_a^2 + 4K_aC_a}}{2}$$

$$[H_3O^+] = -\frac{(1.06\times10^{-3})}{2} + \frac{\sqrt{(1.06\times10^{-3})^2 + 4(1.06\times10^{-3})(1.0\times10^{-2})}}{2}$$
$$= 2.77\times10^{-3}\ \text{M}$$
$$\text{pH} = -\log(2.77\times10^{-3}) = 2.56$$

EXAMPLE 7-13
Calculate pH

- Calculate the pH of a 1 g/100 mL solution of ephedrine sulfate. The molecular weight of the salt is 428.5 g/mol, and Kb for ephedrine base is 2.3 × 10−5.
- (a) The ephedrine sulfate, (BH+)2SO4, dissociates completely into two BH+ cations and one SO42− anion. Thus, the concentration of the weak acid (ephedrine cation) is twice the concentration, Cs, of the salt added.

$$\boxed{[H_3O^+] = \sqrt{K_a C_a}}$$

$K_a = K_w/K_b$

$$C_a = 2C_s = \frac{2 \times 10\ g/L}{428.5\ g/mole} = 4.67 \times 10^{-2}\ M$$

(b)

$$K_a = \frac{1.00 \times 10^{-14}}{2.3 \times 10^{-5}} = 4.35 \times 10^{-10}$$

(c)

$$[H_3O^+] = \sqrt{(4.35 \times 10^{-10}) \times (4.67 \times 10^{-2})} = 4.51 \times 10^{-6}\ M$$

All assumptions are valid. We have

$$pH = -\log(4.51 \times 10^{-6}) = 5.35$$

Solutions Containing Only a Weak Base

- Conc. Of acid (Ca) is zero,
- and [OH−] is generally much greater than [H3O+].

$$[OH^-] = \frac{-K_b + \sqrt{-K_b^{\,2} + 4K_b C_b}}{2}$$

and if Cb is much greater than [OH−], which is generally true for solutions of weak bases,

$$[OH^-] = \sqrt{K_b C_b}$$

EXAMPLE 7-14

Calculate pH

What is the pH of a 0.0033 M solution of cocaine base, which has $K_b = 2.6 \times 10^{-6}$? We have $\boxed{[OH] = \sqrt{K_b C_b}}$

$$[OH^-] = \sqrt{(2.6\times10^{-6})\times(3.3\times10^{-3})}$$
$$= 9.26\times10^{-5}\ M$$

All assumptions are valid. Thus,

$$pOH = -\log(9.26\times10^{-5}) = 4.03$$
$$pH = 14.00 - 4.03 = 9.97$$

Solutions Containing a Single Conjugate Acid–Base Pair

- In a solution composed of a weak acid and a salt (conjugate base) of that acid (e.g., acetic acid and sodium acetate)
- or a weak base and a salt (conjugate acid) of that base (e.g., ephedrine and ephedrine hydrochloride),
- Ca and Cb are generally **much greater** than either [H3O+] or [OH−].

$$[H_3O^+] = \frac{K_a C_a}{C_b}$$

EXAMPLE 7-16

Calculate pH

What is the pH of a solution containing acetic acid 0.3 M and sodium acetate 0.05 M? We write

$$\boxed{[H_3O^+] = \frac{K_a C_a}{C_b}}$$

$$[H_3O^+] = \frac{(1.75 \times 10^{-5}) \times (0.3)}{5.0 \times 10^{-2}}$$
$$= 1.05 \times 10^{-4}\ M$$

All assumptions are valid. Thus,

$$pH = -\log(1.05 \times 10^{-4}) = 3.98$$

EXAMPLE 7-17
Calculate pH

What is the pH of a solution containing ephedrine 0.1 M and ephedrine hydrochloride 0.01 M? Ephedrine has $K_b = 2.3 \times 10^{-5}$; thus, K_a for its conjugate acid is 4.35×10^{-10}.

$$\boxed{[H_3O^+] = \frac{K_a C_a}{C_b}}$$

$$[H_3O^+] = \frac{(4.35 \times 10^{-10}) \times (1.0 \times 10^{-2})}{1.0 \times 10^{-1}}$$
$$= 4.35 \times 10^{-11}$$

All assumptions are valid. Thus,

$$pH = -\log(4.35 \times 10^{-11}) = 10.36$$

Solutions Containing Two Weak Acids

- In systems containing two weak acids, Cb1 and Cb2 are zero,
- For all systems of practical importance, Ca1 and Ca2 are much greater than K1 and K2, so the equation simpli es to

$$[H_3O^+]^2 + [H_3O^+](K_1 + K_2) - (K_1 C_{a1} + K_2 C_{a2}) = 0 \quad (7\text{-}126)$$

If $C_{a1} >> [H_3O+]$ and $C_{a2} >> [H_3O+]$, the equation simplifies to

$$[H_3O^+] = \sqrt{K_1 C_{a1} + K_2 C_{a2}} \quad (7\text{-}127)$$

EXAMPLE 7-21

Calculate pH

What is the pH of a solution containing acetic acid, 0.01 mole/L, and formic acid, 0.001 mole/L? We have $\boxed{[H_3O^+] = \sqrt{K_1 C_{a1} + K_2 C_{a2}}}$

$$[H_3O^+] = \sqrt{(1.75\times10^{-5})(1.0\times10^{-2}) + (1.77\times10^{-4})(1.0\times10^{-3})}$$

$$= 5.93\times10^{-4}\ M$$

$$pH = -\log(5.93\times10^{-4}) = 3.23$$

Solutions Containing a Salt of a Weak Acid and a Weak Base

- The salt of a weak acid and a weak base, such as ammonium acetate, dissociates almost completely in aqueous solution to yield NH4+ and Ac−, the NH4+ is an acid and can be designated as HB1, and the base Ac− can be designated as B2− in equations .
- ***In most instances, however, Cs >> [H3O+],***

$$[H_3O^+] = \sqrt{K_1 K_2}$$

EXAMPLE 7-22

$$\underset{\text{Acid}_1}{NH_4^+} + \underset{\text{Base}_2}{AC^-} \rightleftharpoons \underset{\text{Acid}_2}{HAc} + \underset{\text{Base}_1}{NH_3}$$

Calculate pH

Calculate the pH of a 0.01 M solution of ammonium acetate. The acidity constant for acetic acid is $K_2 = K_a = 1.75 \times 10^{-5}$, and the basicity constant for ammonia is $K_b = 1.74 \times 10^{-5}$. (a) K_1 can be found by dividing K_b for ammonia into K_w:

$$\boxed{[H_3O^+] = \sqrt{K_1 K_2}}$$

$$K_1 = \frac{1.00 \times 10^{-14}}{1.74 \times 10^{-5}} = 5.75 \times 10^{-10}$$

$$[H_3O^+] = \sqrt{(5.75 \times 10^{-10}) \times (1.75 \times 10^{-5})}$$

$$= 1.00 \times 10^{-7} \text{ M}$$

Note that all of the assumptions are valid. We have

$$pH = -\log(1.00 \times 10^{-7}) = 7.00$$

Solutions Containing a Weak Acid and a Weak Base

$$\underset{\text{Acid}_1}{HC_6H_5O_7^{2-}} + \underset{\text{Base}_2}{H_2PO_4^-} \rightleftharpoons \underset{\text{Acid}_2}{H_3PO_4} + \underset{\text{Base}_1}{C_6H_5O_7^{3-}}$$

$$\boxed{[H_3O^+] = \sqrt{K_1 K_2}}$$

EXAMPLE 7-24

Calculate pH

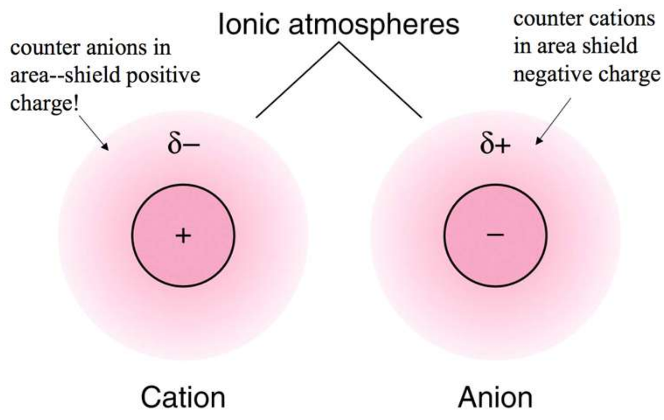
What is the pH of a solution containing NaH_2PO_4 and disodium citrate (disodium hydrogen citrate) $Na_2HC_6H_5O_7$, both in a concentration of 0.01 M? The third acidity constant for $HC_6H_5O_7^{2-}$ is 4.0×10^{-7}, whereas the first acidity constant for phosphoric acid is 7.5×10^{-3}. We have

$$[H_3O^+] = \sqrt{(4.0 \times 10^{-7}) \times (7.5 \times 10^{-3})}$$
$$= 5.48 \times 10^{-5}\ \text{M}$$

All assumptions are valid. We find

$$\text{pH} = -\log(5.48 \times 10^{-5}) = 4.26$$

Ionic strength



Ionic strength

- ionic strength---I or μ ---a measure of the total ion concentration in solution----but ions with more charge are counted more due to stronger electrostatic interactions with other ions (I.e., can influence the increase "ionic atmosphere" greater than singly charged ions)

$$\mu = \frac{1}{2}\sum_i c_i z_i^2$$

where c_i is conc. of i^{th} species and z_i is the charge on i^{th} species

What is ionic strength of 0.01 M NaCl solution?

$\mu = 1/2\,([Na^+]z_{Na}^2 + [Cl^-]z_{Cl}^2) = 1/2\,(0.01\,(1)^2 + 0.01(-1)^2) = 0.01$ M

What is ionic strength of 0.01 M Na_2SO_4 solution?

$\mu = 1/2([Na^+]z_{Na}^2 + [SO_4^{-2}]z_{SO4}^2) = 1/2(0.02\,(1)^2 + 0.01\,(-2)^2) = 0.03$ M

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Effect of Ionic Strength on Acidity Constants

$$HB + H_2O \rightleftharpoons H_3O^+ + B$$

$$K = \frac{\alpha_{H_3O} + \alpha_B}{\alpha_{HB}} = \frac{[H_3O^+][B]}{[HB]} \cdot \frac{\gamma H_3O^+ \gamma_B}{\gamma_{HB}}$$

For monotropic molecule

$$\boxed{pK' = pK + \frac{0.51(2Z - 1)\sqrt{\mu}}{1 + \sqrt{\mu}}}$$

For zwitterion molecule

$$\boxed{pK_1' = pK_1 + \frac{0.51\sqrt{\mu}}{1+\sqrt{\mu}} - K_r\mu}$$

$$\boxed{pK_2' = pK_2 - \frac{0.51\sqrt{\mu}}{1+\sqrt{\mu}} + K_r\mu}$$

K_r=Salting in constant=0.32 for amino acid in water

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EXAMPLE 7–27

Calculate pH

Calculate the pH of a 0.01 M solution of acetic acid to which enough KCl had been added to give an ionic strength of 0.1 M at 25°C. The pK_a for acetic acid is 4.76.

(a) $\boxed{pK' = pK + \frac{0.51(2Z - 1)\sqrt{\mu}}{1 + \sqrt{\mu}}}$

$$pK'_a = 4.76 - \frac{0.51\sqrt{0.10}}{1 + \sqrt{0.10}}$$

$$= 4.76 - 0.12 = 4.64$$

(b) Taking logarithms of equation (7–99) gives

$$[H_3O^+] = K_a \frac{(C_a - [H_3O^+] + [OH^-])}{(C_b + [H_3O^+] - [OH^-])}$$

$$pH = \frac{1}{2}(pK'_a - \log C_a)$$

in which we now write pK_a as pK'_a:

$$pH = \frac{1}{2}(4.64 + 2.00) = 3.32$$

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EXAMPLE 7–28

Calculate pH

Calculate the pH of a 10^{-3} M solution of glycine at an ionic strength of 0.10 at 25°C. The pK_a values for glycine are $pK_1 = 2.35$ and $pK_2 = 9.78$.

(a)
$$pK'_1 = 2.35 + \frac{0.51\sqrt{0.10}}{1 + \sqrt{0.10}} - 0.32(0.10)$$

$$= 2.35 + 0.12 - 0.03 = 2.44$$

(b)
$$pK'_2 = 9.78 - \frac{0.51\sqrt{0.10}}{1 + \sqrt{0.10}} + 0.32(0.10)$$

$$= 9.78 - 0.12 + 0.03 = 9.69$$

(c) Taking logarithms of equation (7–118) gives

$$pH = \frac{1}{2}(pK_1 + pK_2)$$

$$\boxed{[H_3O^+] = \sqrt{K_1K_2}}$$

$$= \frac{1}{2}(2.44 + 9.69) = 6.07$$

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Thanks for your attention