

Chem 103

Lecture 3b
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Acids and Bases1

Last time

- Chapter 14 – intro to solutions
- Solution concentrations
- Henry's Law
- Colligative properties

Today

- Finish Colligative properties
- Acids and Bases

Colligative properties

Colligative properties = physical properties of solutions which depend upon the concentration of solute particles.

- 4 main examples:
- a) Boiling point elevation
 - b) Freezing point depression
 - c) Vapor Pressure lowering
 - d) Osmotic pressure

Osmotic pressure

$\pi = iMRT$ (across semipermeable membrane)
Where π = osmotic pressure
 i = "van't Hoff factor", a correction factor to account for actual # particles in solution
 M = molarity of solute
 $R = 0.0821 \text{ atm}\cdot\text{L}/\text{mol}\cdot\text{K}$
 T = temperature in K

Osmotic pressure problem

The equation is: $\pi = cRTi$ where c = moles/L (molarity)

(can be derived from $pV=nRT \Rightarrow p=(n/V)RT$)

Example. A solution prepared by adding 50. g of solute to make 1.0 L solution at 300 K has $\pi = .821 \text{ atm}$. What is the MW of the solute (assuming it is a nonelectrolyte)?

$$\pi = cRTi = \frac{\text{g}_{\text{solute}}/\text{MW}_{\text{solute}}}{V_{\text{solution}}} RTi$$

$$\text{or, } \text{MW} = \frac{\text{g}_{\text{solute}}}{\pi V_{\text{solution}}} RTi$$

$$\text{MW} = \frac{50. \text{g}}{(.821 \text{ atm})(1.0 \text{ L})} (0.0821 \frac{\text{atm}\cdot\text{L}}{\text{mol}\cdot\text{K}})(300 \text{ K})(1)$$

$$\text{MW} = 1.50 \times 10^3 \text{ g/mol}$$

Red blood cells and π

Red blood cells (RBC's) are "semipermeable bags", which must maintain the same concentration within and without, or else! (i.e. solution surrounding it must be *isotonic*)

Water leaves cell: crenation

If solution is more concentrated (*hypertonic*) than the internal concentration, what happens?

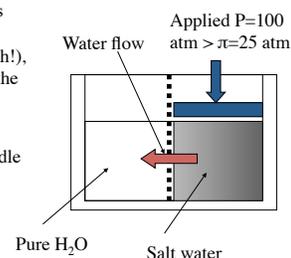
If solution is less concentrated (*hypotonic*) than the internal concentration, what happens?

Water enters cell: hemolysis

Reverse osmosis: desalination

If the applied pressure is high enough (and the membrane robust enough!), it is possible to reverse the flow of solvent.

Example: desalination plants - common in middle east, Florida,...



Definition of Acid and Base

Arrhenius's definition:

acid = proton donor,

examples: HCl, HNO₃

base = hydroxide donor

examples: NaOH, Mg(OH)₂

Definition of Acid and Base: Arrhenius' definition

Arrhenius's definition: $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$

acid = proton donor,

base = hydroxide donor,

examples: HCl vs NaOH

$\text{NaOH} \rightarrow \text{Na}^+ + \text{OH}^-$

"H⁺" = "H₃O⁺"

=hydronium ion

Bronsted-Lowry definition

Bronsted-Lowry's is a broader definition:

acid = proton donor,

base = *proton acceptor*.

example: HCl vs NH₃

HCl donates a proton: $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$

NH₃ accepts a proton:

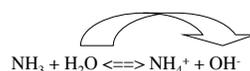
$\text{NH}_3 + \text{H}^+ \rightarrow \text{NH}_4^+$

B-L def: Acid-base conjugates

Example of ammonia illustrates Bronsted-Lowry definition and also the concept of *conjugate acids and bases*

NH₃ is conjugate base of NH₄⁺ while

conjugates NH₄⁺ is conjugate acid of NH₃



H₂O is the conjugate acid of OH⁻ while OH⁻ is the conjugate base of H₂O

conjugates

Role of Water

Water is a very important solvent. It can donate or accept protons: it is therefore *AMPHIPROTIC*.

As an acid: $\text{H}_2\text{O} \rightarrow \text{H}^+ + \text{OH}^-$

As a base: $\text{H}^+ + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+$

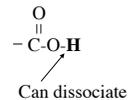
Does water act as a base or an acid when HCl is added to it?

Carboxylic acids and amines

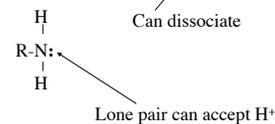
Many weak acids and weak bases are organic compounds - i.e. based on carbon.

Most common acidic group:

carboxyl group

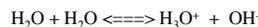


Most common basic group: amine group



Self-ionization of water

Even pure water can slightly self ionize as follows:



For pure water: $[\text{H}^+] = 1.0 \times 10^{-7} \text{M}$ and $[\text{OH}^-] = 1 \times 10^{-7}$

So, $K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = (1.0 \times 10^{-7})^2 = 1.0 \times 10^{-14}$

Note that K_w is a constant (changes slightly with temp.)

So for acidic solutions:

$[\text{H}_3\text{O}^+]$ is high due to presence of acid, and $[\text{OH}^-]$ is low

But $[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$ still.

Strengths of acids and bases

Acids and Bases are categorized into two general types:

Strong and weak. Refers to the tendency to donate or accept H^+

Strong acids and bases dissociate 100%: HCl, HNO_3 ; NaOH, KOH

Weak acids and bases do not dissociate 100%: CH_3COOH , HF, NH_3

Among weak acids and bases, the rule of thumb is:

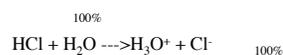
"the stronger the conjugate, the weaker the acid or base"

Strengths of acids & bases:

Conjugate acids		Conjugate bases
HCl	↑ Increasing strength	Cl^-
HAc		Ac^-
H_2O		OH^-
		↓ Increasing strength

Dissociation of strong acids in water

Strong acids dissociate 100% in water. (Classic example: HCl)



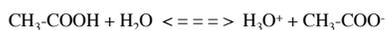
Other examples: $\text{HNO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{NO}_3^-$

HBr, HI, H_2SO_4 (remember all these strong acids)

Often we just write: $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$; $\text{HNO}_3 \rightarrow \text{H}^+ + \text{NO}_3^-$

Weak acid solution: K_a equilibrium

Weak acids don't ionize 100% (example: acetic acid)



General eq'n: $\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}^-$

" K_a equilibrium"; K_a is "acid dissociation constant".

Expression for K_a : $K_a = [\text{H}_3\text{O}^+][\text{A}^-] / [\text{HA}]$

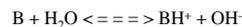
The *larger* the K_a , the *stronger* the weak acid.

E.g.: for $\text{CH}_3\text{CO}_2\text{H}$, $K_a = 1.8 \times 10^{-5}$

for NH_4^+ $K_a = 5.6 \times 10^{-10}$. $\text{CH}_3\text{CO}_2\text{H}$ is the stronger acid.

Weak base solution: K_b equilibrium

For weak bases: " K_b equilibrium". K_b = "base ionization constant". General equation:



$$K_b = [\text{BH}^+][\text{OH}^-] / [\text{B}]$$

The greater K_b is, the stronger the base.

E.g.: for NO_2^- , $K_b = 2.2 \times 10^{-11}$ for NH_3 , $K_b = 1.8 \times 10^{-5}$

NH_3 is the stronger base.

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

Be able to prove this:

Since $K_a = [\text{H}_3\text{O}^+][\text{A}^-] / [\text{HA}]$ for acid HA.

For its conjugate base, A^- , it is: $K_b = [\text{HA}][\text{OH}^-] / [\text{A}^-]$

So if multiply K_a of weak acid by K_b of its conjugate, then:

$$K_a K_b = ([\text{H}_3\text{O}^+][\text{A}^-] / [\text{HA}]) ([\text{HA}][\text{OH}^-] / [\text{A}^-])$$

$$K_a K_b = [\text{H}_3\text{O}^+][\text{OH}^-] = K_w = 1.0 \times 10^{-14}$$

This refers to conjugate pairs.

pH

Because of wide range of $[\text{H}_3\text{O}^+]$, it's convenient to express concentration levels of $[\text{H}^+]$ by its exponent, using logarithmic scale. This is the "pH scale".

Definition of pH: **pH = $-\log[\text{H}^+] = -\log[\text{H}_3\text{O}^+]$**

pH of pure water: $[\text{H}_3\text{O}^+] = 1.0 \times 10^{-7} \text{M}$

$\text{pH} = -\log\{[\text{H}_3\text{O}^+]\} = -\log(1.0 \times 10^{-7} \text{M}) = 7.00$

This is the *neutral* pH

pH of acidic solutions

If for acidic solution $[\text{H}_3\text{O}^+] = 2.0 \times 10^{-4}$. What's the pH?

$\text{pH} = -\log(2.0 \times 10^{-4} \text{M}) = 3.70$

In general, $\text{pH} < 7.00$ is acidic...

And $\text{pH} > 7.00$ is basic.

Calculating $[\text{H}^+]$ given the pH: **$[\text{H}^+] = 10^{-\text{pH}}$**

pOH

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

Since: $[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$

$-\log\{[\text{H}_3\text{O}^+][\text{OH}^-]\} = -\log\{1.0 \times 10^{-14}\}$

$-\log[\text{H}_3\text{O}^+] - \log[\text{OH}^-] = 14.00$

Or, $\text{pH} + \text{pOH} = 14.00$

(so if you know pH, you can know pOH)

Summary

$$[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+] \quad \text{and} \quad \text{pOH} = -\log[\text{OH}^-]$$

$$\text{pH} + \text{pOH} = 14.00$$

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} \quad K_b = \frac{[\text{HA}][\text{OH}^-]}{[\text{A}^-]}$$

$$K_a K_b = 1.0 \times 10^{-14}$$

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

Next time: buffers