

Chem 103 lecture 2a

Finish osmotic pressure and water treatment

Acid/Base - definitions

Role of water Strengths of acids and bases

K_a and K_b

Osmotic pressure problem

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

(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)?

$$\pi = cRTi = \frac{g_{\text{solute}}/MW_{\text{solute}}}{V_{\text{solution}}} RTi$$

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

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

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

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?

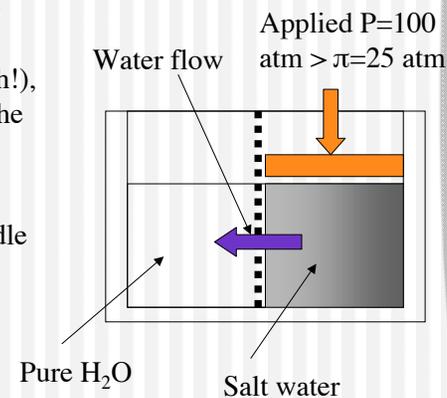
If solution is less concentrated (*hypotonic*) than the internal concentration, what happens?

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,...



Colloids

Colloid = dispersed phase (large molecs, 2-2000nm) + continuous phase (solvent)

=often scatter light (Tyndall effect)

Examples:

Aerosols (liquid or solids and gas) like fog, clouds

Foam (gas in liquid) like whipped cream

Emulsion (liquids in liquid) like milk

Gels (solids in liquids) like butter

Surfactants

Surfactants = molecules with hydrophobic and hydrophilic parts

e.g. sodium stearate: $\text{Na}^+ \text{O-C(=O)-(CH}_2\text{)}_{16}\text{-CH}_3$

Soaps are made by saponification reaction of triglycerides

Triglycerides + 3 NaOH \rightarrow Na(fatty acid) + glycerine

Soaps + grease + water = an emulsion...

Municipal water treatment

Only 3% of all the planet's water is fresh water.

Only 0.008% is found in reservoirs, rivers, lakes

0.77% of water is ground water.

LA is in 3rd year of water shortage: June 1 start of mandatory water conservation! Water treatment is very important.

Waste water intake \rightarrow coarse screen \rightarrow settling tanks \rightarrow sand filter

\rightarrow aeration \rightarrow Cl or O₃ addition \rightarrow storage tank \rightarrow consumer.

“Water hardness”: high concentration of Ca²⁺, Mg²⁺, Fe²⁺, or Mn²⁺

Definition of Acid and Base

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

acid = proton donor,

base = hydroxide donor,

examples: HCl vs NaOH

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

A broader definition: Bronsted-Lowry:

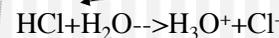
acid = proton donor,

base = *proton acceptor*.

example: HCl vs NH_3

“H⁺” = “H₃O⁺”

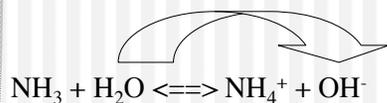
=hydronium ion



B-L def: Acid-base conjugates

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

NH_3 is conjugate base of NH_4^+ while
conjugates NH_4^+ is conjugate acid of NH_3



H_2O is the conjugate acid of OH^- while
 OH^- is the conjugate base of H_2O

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?

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

HCl

HAc

H₂O

↑
Increasing strength

Conjugate bases

Cl⁻

Ac⁻

OH⁻

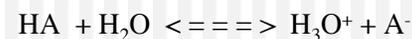
↓
Increasing strength

Dissociation of acids in water

Strong acids dissociate 100% in water, we normally assume



Weak acids ionize < 100% (for example acetic acid):



Called “K_a equilibrium”. K_a is “acid dissociation constant”.

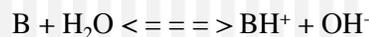
What is the expression for K_a ? $K_a = [\text{H}_3\text{O}^+][\text{A}^-] / [\text{HA}]$

The larger the K_a, the stronger the weak acid.

Usually K_a very small, ex : K_a for acetic acid = 1.8x10⁻⁵

Weak base equilibrium

For weak bases: “K_b equilibrium”. K_b = “base ionization constant:



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

The greater K_b is, the stronger the base.

Usually K_b very small, example: NO₂⁻ has K_b = 2.2x10⁻¹¹

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

Note: K_a = [H₃O⁺][A⁻]/[HA] for acid HA.

For it's conjugate, A⁻, it is: K_b = [HA][OH⁻] / [A⁻]

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

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

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

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

For example: what's the pH of pure water?

Note: 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 you have an acidic solution with $[\text{H}_3\text{O}^+] = 2.0 \times 10^{-4}\text{M}$.
What's the pH?

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

In general, if $\text{pH} < 7.00$ we say solution is acidic...

If $\text{pH} > 7.00$ the solution is basic.

What is $[\text{H}^+]$ if you are given the pH?

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

pOH

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

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

(so if you know $[\text{H}_3\text{O}^+]$, you can know $[\text{OH}^-]$)

take -log of both sides:

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

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

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

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

Summary

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

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

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

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

$K_a K_b = 14.00$

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

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