Review of Elementary Acid – Base Chemistry: Extension and Application of the Concepts to Organic Reactivity

For 2nd and 3rd Year Organic and Biochemistry Students

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Review: Brønsted – Lowry Definitions of Acids and Bases

- According to the Brønsted – Lowry definitions:
  - *An acid* is a species that *increases the proton concentration* of a solution.
  - *A base* is a species that *increases the hydroxide concentration* of a solution.

- Examples of Brønsted – Lowry acids:
  - HCl:

$$\mathrm{HCl(g)} \xrightarrow{\mathrm{H_2O}} \mathrm{H^+(aq) + Cl^-}(aq)$$

A molecule of hydrogen chloride gas, $HCl_{(g)}$, interacts with water by *dissociating* into a proton and a chloride ion each surrounded by water, $H^+_{(aq)}$ and $Cl^-_{(aq)}$ respectively. This increases the proton concentration of the solution: *HCl(g) is an acid*.

  - Acetic acid:

$$\mathrm{CH_3COOH(aq)} \rightleftharpoons \mathrm{H^+(aq) + CH_3COO^-}(aq)$$

A molecule of acetic acid dissolved in water, $CH_3COOH_{(aq)}$, *dissociates* into a proton and an acetate ion each surrounded by water $H^+_{(aq)}$, $CH_3COO^-_{(aq)}$ respectively. This increases the proton concentration of the solution: *acetic acid is an acid*.

Review: Brønsted – Lowry Definitions of Acids and Bases

- According to the Brønsted – Lowry definitions:
  - *An acid* is a species that *increases the proton concentration* of a solution.
  - *A base* is a species that *increases the hydroxide concentration* of a solution.

- Examples of Brønsted – Lowry bases:
  - NaOH:

$$\mathrm{NaOH(s)} \xrightarrow{\mathrm{H_2O}} \mathrm{Na^+(aq) + OH^-}(aq)$$

Solid sodium hydroxide, $NaOH_{(s)}$ dissolves in water by *dissociating* into a sodium ion and an hydroxide anion, $Na^+_{(aq)}$ and $OH^-_{(aq)}$ respectively. This increases the hydroxide concentration of the solution: *sodium hydroxide is a base*.

  - Ammonia:

$$\mathrm{NH_3(aq)} \overset{\mathrm{H_2O}}{\rightleftharpoons} \mathrm{NH_4^+\ (aq) + OH^-}(aq)$$

"Ammonia dissolved in water, $NH_{3(aq)}$ reacts with water by *abstracting* protons from water, thereby creating ammonium cations and hydroxide anions, $NH_4{}^+_{(aq)}$ and $OH^-_{(aq)}$ respectively. This increases the hydroxide concentration of the solution: *ammonia is a base.*

Review: Strong Acids and Bases

- *Strong acids and bases 'dissociate fully' in aqueous solution*. The reaction 'goes to completion.'
  - HCl:

$$\mathrm{HCl(g)} \xrightarrow{\mathrm{H_2O}} \mathrm{H^+(aq)} + \mathrm{Cl^-}(aq)$$

  - NaOH:

$$\mathrm{NaOH(s)} \xrightarrow{\mathrm{H_2O}} \mathrm{Na^+(aq)} + \mathrm{OH^-}(aq)$$

Notice the arrows!

Unidirectional arrows in chemical equations describe processes in which the species on the left is (are) fully transformed into the species on the right.

Review: Weak Acids and Bases

- In contrast, *the dissociation of weak acids and bases are equilibrium processes*, and chemical equations describing them employ *double harpoon arrows.* Reactions involving weak acids and bases do not 'go to completion,' instead they reach *equilibrium,* a point at which an *unchanging* concentration of *each of the species* in the chemical equation occurs in solution.

  - Acetic acid:

$$CH_3COOH(aq) \overset{H_2O}{\rightleftharpoons} H^+(aq) + CH_3COO^-(aq)$$

  - Ammonia:

$$NH_3(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Notice the double harpoon arrows!

Review: Weak Acid and Base Equilibrium Constants

Weak acid and base equilibrium constants, K_A and K_B, are constructed in the normal way: *

- Acetic acid:

$$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$$

$$K_A = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$$

- Ammonia:

$$NH_3(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

$$K_B = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

**: It is beyond the scope of this presentation to review (a) the construction of equilibrium constants or (b) to explain the non-occurrence of water concentration terms in them. For these concepts, consult any standard general, organic, or biochemistry text.*

Equilibrium Constants Facilitate the Determination of Equilibrium Concentrations (I)

Suppose that for acetic acid, K_A = 1.74 x 10^{-5}. If a solution is prepared such that 1.00 mol / L of acetic acid is dissolved in 1.00 L of water, determine the equilibrium concentrations of acetic acid (CH_3COOH), protons (H^+) and acetate anions (CH3COO$^-$).

	Protons (mol/L)	Acetic acid (mol/L)	Acetate anion (mol/L)
Initial Concentrations	0.00	1.00	0.00
Concentration Changes	+x	-x	+x
Equilibrium Concentrations	0.00 + x	1.00 - x	0.00 + x

$$K_A = \frac{[\mathrm{H^+}][\mathrm{CH_3COO^-}]}{[\mathrm{CH_3COOH}]}$$

Substituting the equilibrium constant and equilibrium concentrations, into algebraic expression for the equilibrium constant provides:

$$1.74 \times 10^{-5} = \frac{[x][x]}{[1.00 - x]}$$

Equilibrium Constants Facilitate the Determination of Equilibrium Concentrations (II)

$$1.74 \text{ x } 10^{-5} = \frac{[x][x]}{[1.00 - x]}$$

The above is a quadratic equation and the value of x may be computed as the *positive* root of the equation using the quadratic formula; negative roots (corresponding to negative concentrations) are meaningless.

The quadratic formula provides:

$$\text{x} = 0.00295 \quad \frac{\text{mol}}{\text{L}}$$

Substituting this value back into the 'Equilibrium Concentrations' equations from the previous table allows computation of the equilibrium concentrations of protons, acetic acid molecules, and acetate anions. These are, respectively, 0.00295 mol/L, 0.997 mol/L, and 0.00295 mol/L.

Note to students: if you do not know the quadratic formula or how to use it, learn it!
If you need help – come see me.

Review: Conjugate Acid / Base Pairs

- The dissociation of a weak acid results in the liberation of a proton and a *conjugate base*.
  - *The acetate anion is the conjugate base of acetic acid:*

$$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$$

$$K_A = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$$

- The hydrolysis of water by a weak base results in the liberation of an hydroxide anion and a *conjugate acid*.
- *The conjugate base of a weak acid is a weak base:*

$$CH_3COO^-(aq) + H_2O \rightleftharpoons CH_3COOH(aq) + OH^-(aq)$$

$$K_B = \frac{[CH_3COOH][OH^-]}{[CH_3COO^-]}$$

Review: Water is Both a Weak Acid and a Weak Base

$$H_3CCOOH \rightleftharpoons H^+ + H_3CCOO^-$$

$$H_3CCOO^- + H_2O \rightleftharpoons H_3CCOOH + OH^-$$

$$H_3CCOOH + H_3CCOO^- + H_2O \rightleftharpoons H^+ + H_3CCOO^- + H_3CCOOH + OH^-$$

$$H_2O \rightleftharpoons H^+ + OH^-$$

Therefore:

$$K_A K_B = K_W$$

$$K_W = 1 \times 10^{-14}$$

$$K_W = [H^+][OH^-]$$

From which, of course, we get the *pH* (and *pOH*) scales.*

** For pH, pOH, the addition of chemical equations and the corresponding manipulation of their equilibrium constants, consult any standard general chemistry text or come to me for assistance.*

Review: The Reaction of a Weak Acid with a Base Generates a Salt and Water

$$CH_3COOH(aq) + NaOH(aq) \rightleftharpoons CH_3COONa(aq) + H_2O$$

Review: The Reaction of a Weak Acid with a Base Generates a Salt and Water

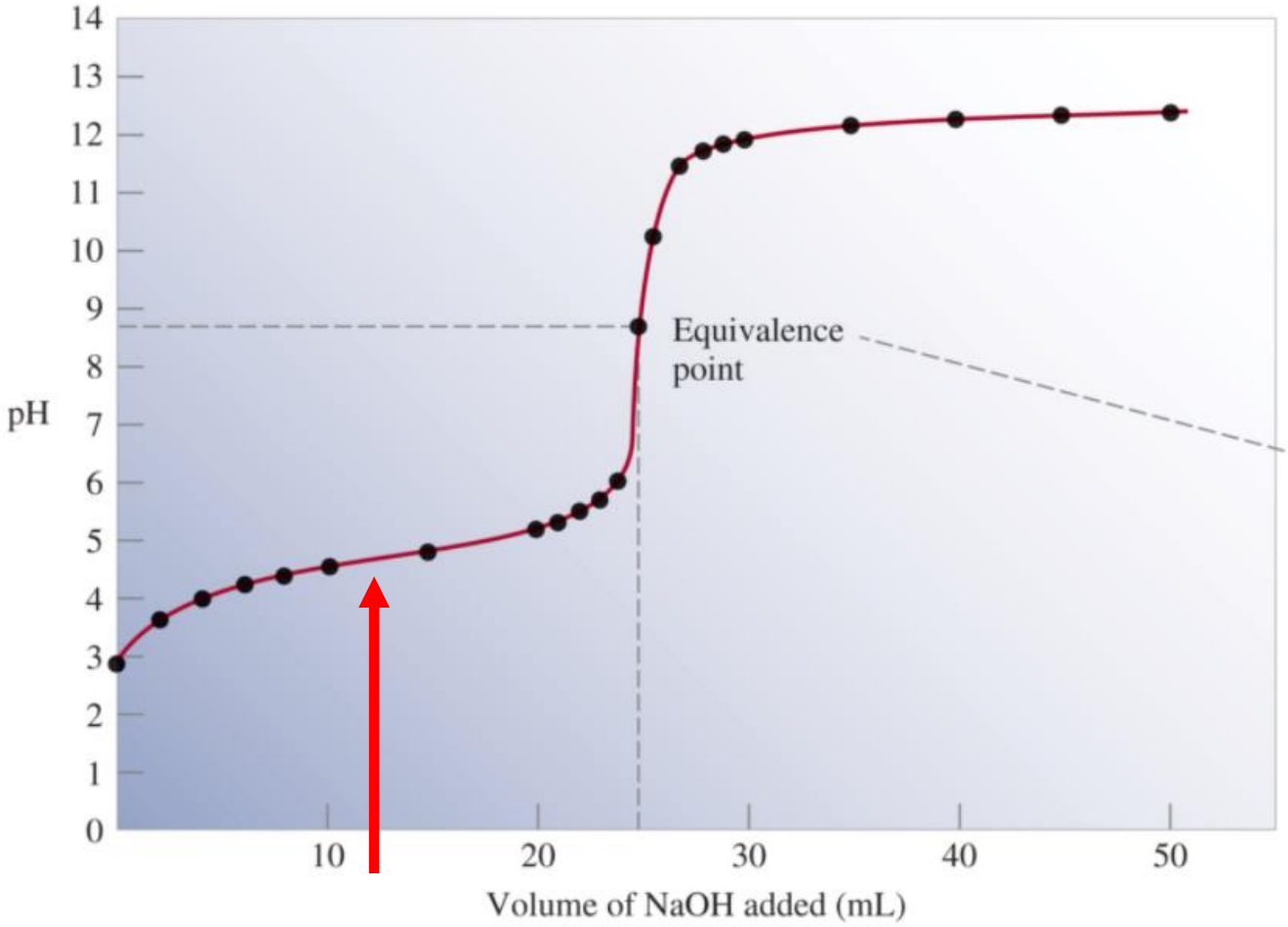
$$\mathrm{CH_3COOH(aq) + NaOH(a}q\mathrm{)} \rightleftharpoons \mathrm{CH_3COONa(aq)} + \mathrm{H_2O}$$

Now:

What if instead of carrying out this reaction quantitatively, it were instead carried out only partially?

That is, instead of reacting an equimolar amount of base with the weak acid, only a fractional amount, say half the molar amount, of the base were added?

Titration of a Weak Acid by a Strong Base



Volume NaOH added (mL)	pH
0.0	2.87
5.0	4.14
10.0	4.57
15.0	4.92
20.0	5.35
22.0	5.61
24.0	6.12
25.0	8.72
26.0	10.29
28.0	11.75
30.0	11.96
35.0	12.22
40.0	12.36
45.0	12.46
50.0	12.52

Note: the *gradual* change in pH as a function of base added.

At each point in the titration (before the equivalence point) the solution consists of a mixture of acid and conjugate base forms. This kind of mixture is a *buffer*.

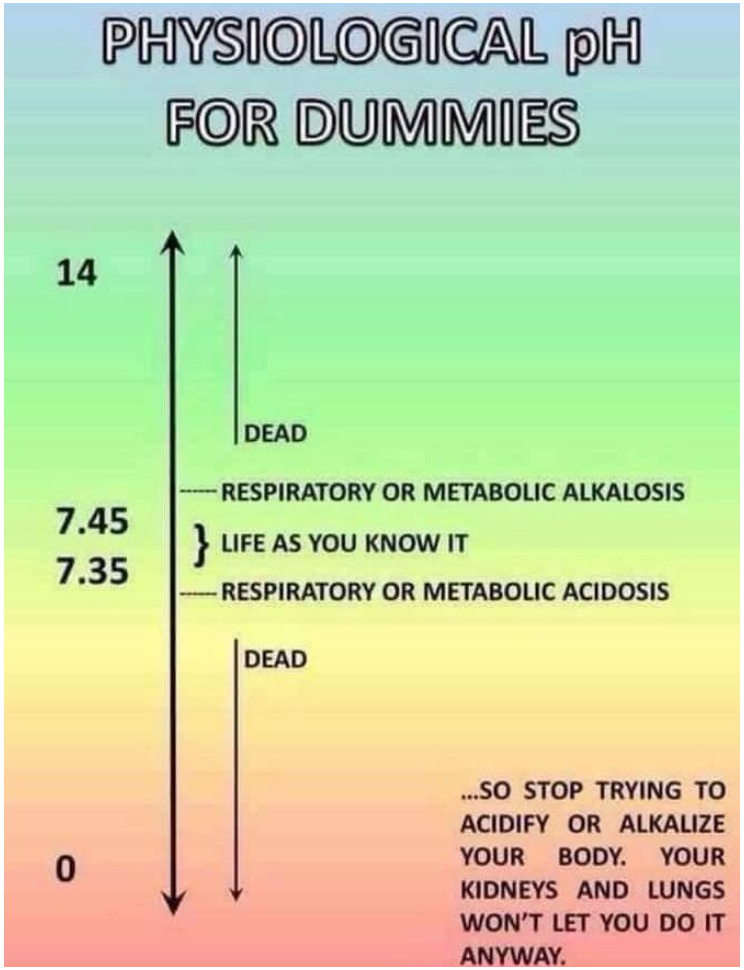
Buffers: Solutions Which are Mixtures of Acid (or base) and Conjugate Base (or acid) Forms

Buffers resist changes in pH.

Buffers are essential.

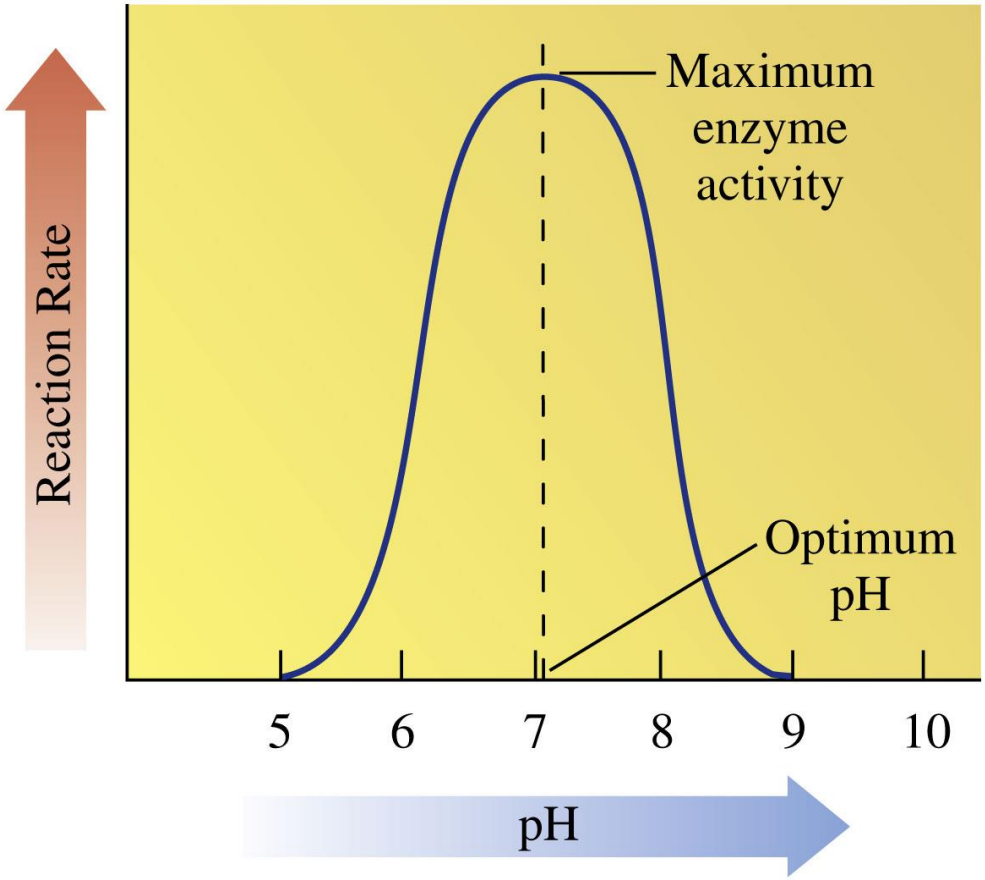
- For health and disease.
  - Example: the blood is a carbonic acid / bicarbonate buffer system which maintains pH in the vicinity of 7.35 – 7.45.

$$H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$$



Why Are Buffers Important in Biochemistry?
pH Affects Enzymatic Activity:

- Much of enzymatic catalysis involves weak acid / weak base chemistry.
  - In other words, proton transfers play a big role in enzymatic catalysis.
  - Therefore, *pH* dramatically affects enzymatic activity.
  - Almost all enzymes only carry out their reactions in a narrow range of *pH*.



Can anyone think of *another* reason it is preferable to study proteins in buffered solutions?

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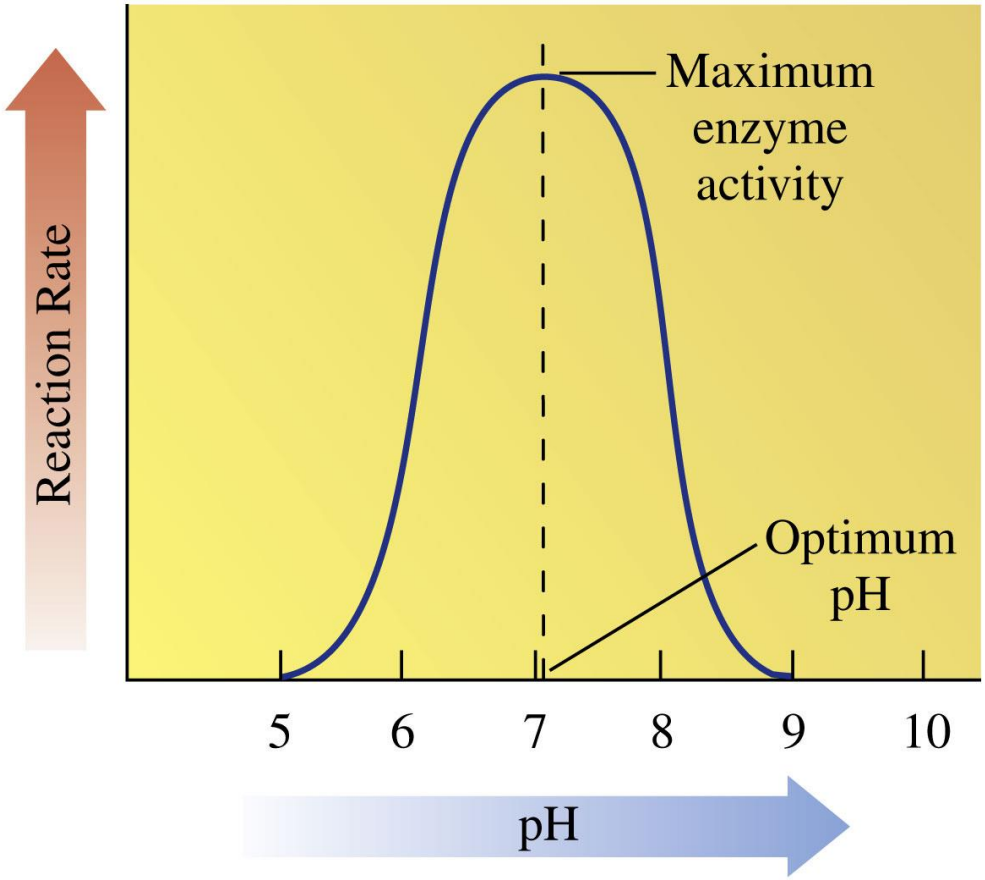

Can anyone think of *another* reason it is preferable to study proteins in buffered solutions?

Proteins are built of amino acids which themselves contain acidic and basic groups.

The acid-base chemistry of amino acid side chains would disrupt the *pH* of unbuffered protein solutions.

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The Henderson – Hasselbalch Equation

An equation relating pH and the concentrations of acids (bases) and conjugate bases (acids):

$$HA \rightleftharpoons H^+ + A^-$$

$$K_A = \frac{[H^+][A^-]}{[HA]}$$

Solving for $[H^+]$:

$$[H^+] = \frac{K_A[HA]}{[A^-]}$$

Taking negative logarithms of both sides and rearranging:

$$-\text{Log}[H^+] = -\text{Log}\, K_A - \text{Log}\frac{[HA]}{[A^-]}$$

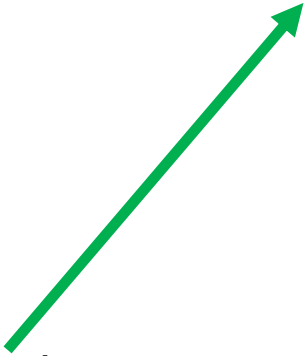
$$pH = pK_A + \text{Log}\frac{[A^-]}{[HA]} *$$

* Recall that pN = -Log(N)

Lewis Acidity and Basicity: More General Definitions of the Concepts

Recalling the Bronsted definition:

$$HA \rightleftharpoons H^+ + A^-$$



What is this?
(Literally.)

Lewis Acidity and Basicity: More General Definitions of the Concepts

Recalling the Bronsted definition:

$$HA \rightleftharpoons H^+ + A^-$$

What is this? It is literally a proton.

What is a (literal) proton missing that *all* other atoms have?

One or more electrons.

A Lewis acid is a species that is *electron deficient:*
or which is more usually described as "an electron pair acceptor."

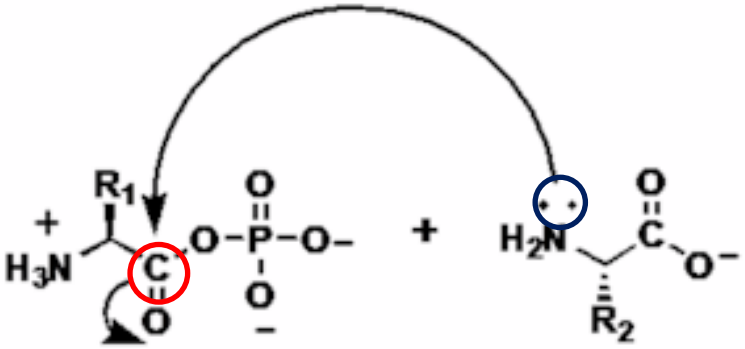
Conversely, a Lewis Base is a species that is *electron rich: or which is more usually described "as an electron pair donor."*

Note that the Brønsted-Lowry and Lewis definitions are *entirely consistent with each other*. They are two different ways of looking at the same phenomenon, but the Lewis understanding is more general.

Why Lewis Acids and Bases?

The concepts of Lewis acidity and basicity provide a framework for understanding much of organic chemical reactivity.

Consider the peptide bond formation that takes place on the ribosome.



- The amino group of the last amino acid added to the growing peptide chain bears a lone pair of electrons, circled in blue.
- This means that it is *electron rich*, one of the definitions of a *Lewis base*.

- The carbonyl carbon of the acylphosphate group on the incoming amino acid, circled in red, is highly *activated*, as in 'primed to react.'
- Involved in a double bond to oxygen and attached to the strongly electron-withdrawing phosphate group, this carbon is very *electron-deficient*. It is a *Lewis acid*.