Classical approach to Acid-Base Balance understanding

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Praise of the bicarbonate

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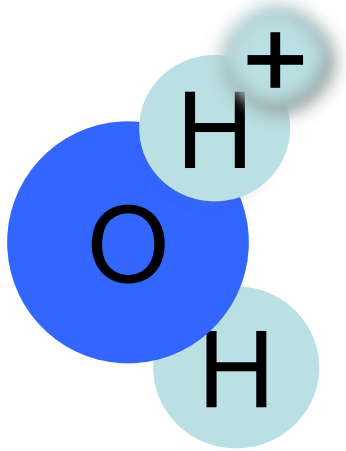
The talk

- What are acids and bases?
- What is bicarbonate?
- What are buffers?
- Buffer base and Base excess
- What causes acidosis? – principle of electroneutrality and Anion-Gap

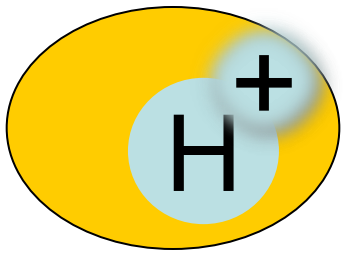
What are acids and bases?

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

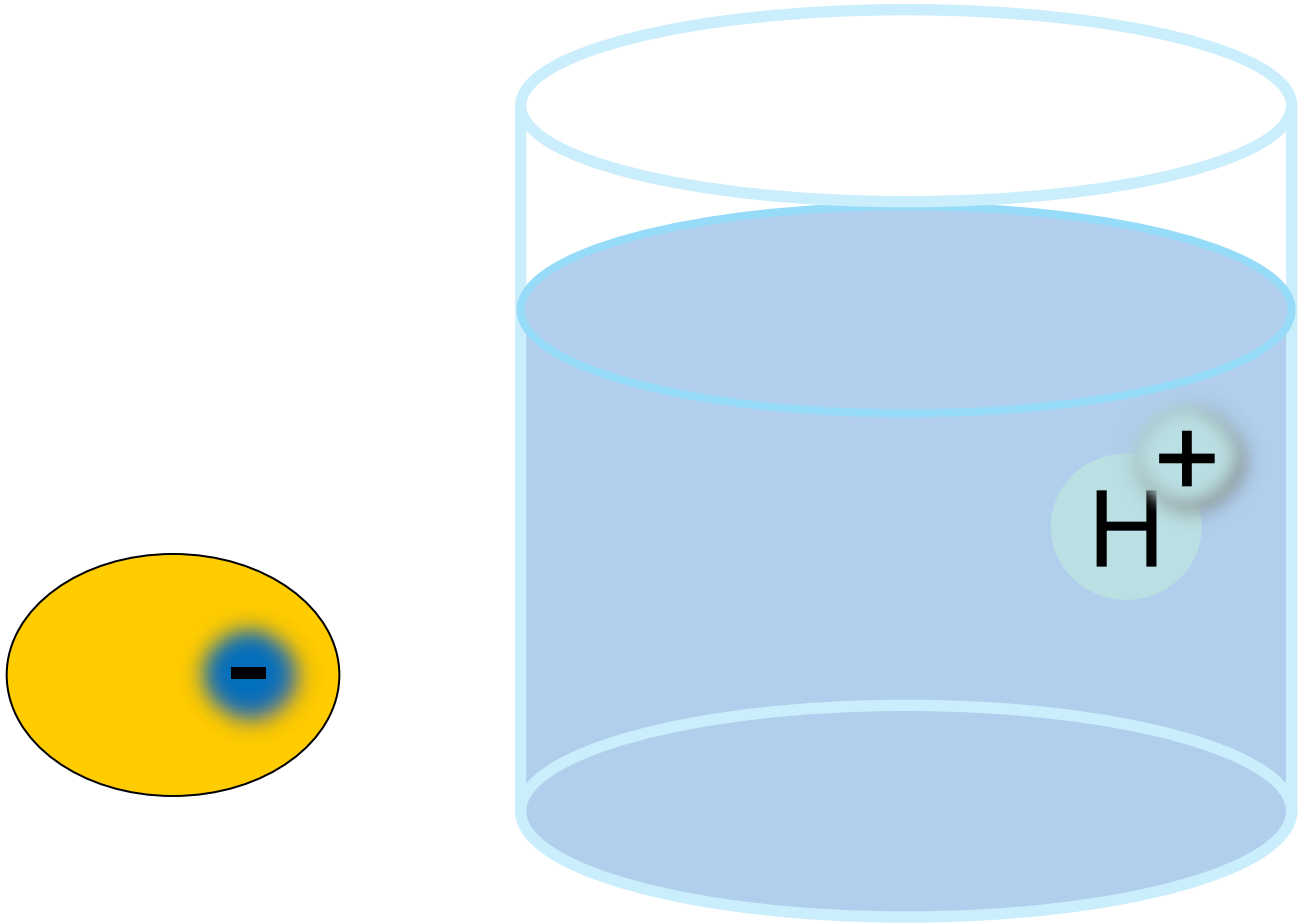


$[H^+] = 40\ nmol/l$

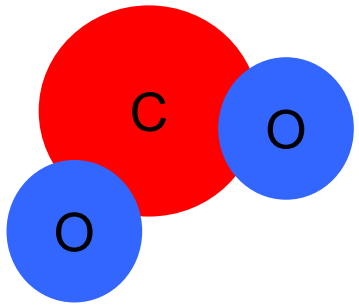
Acid – proton donor



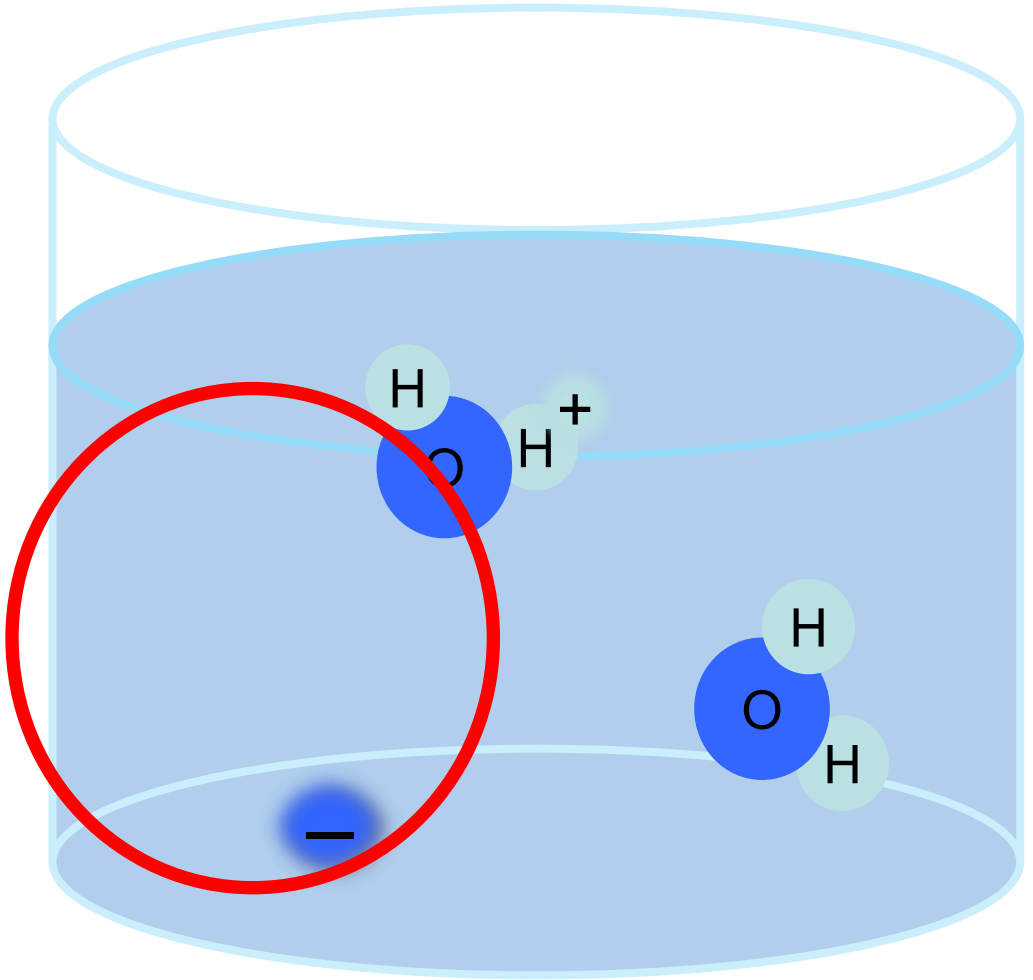
Base – proton acceptor



What is bicarbonate?

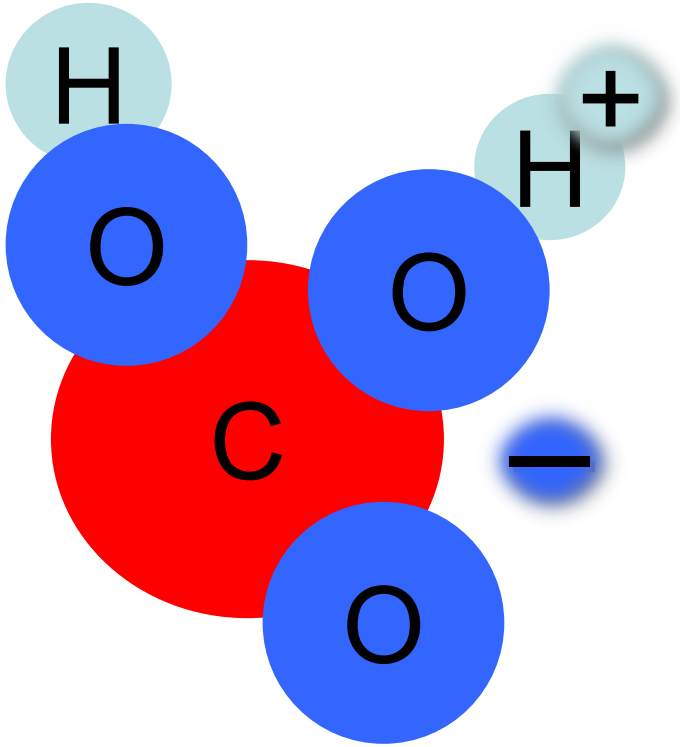


$$CO_2 + H_2O \rightarrow H_2CO_3 \rightarrow H^+ + HCO_3^-$$



Carbonic acid

$H_2CO_3^-$

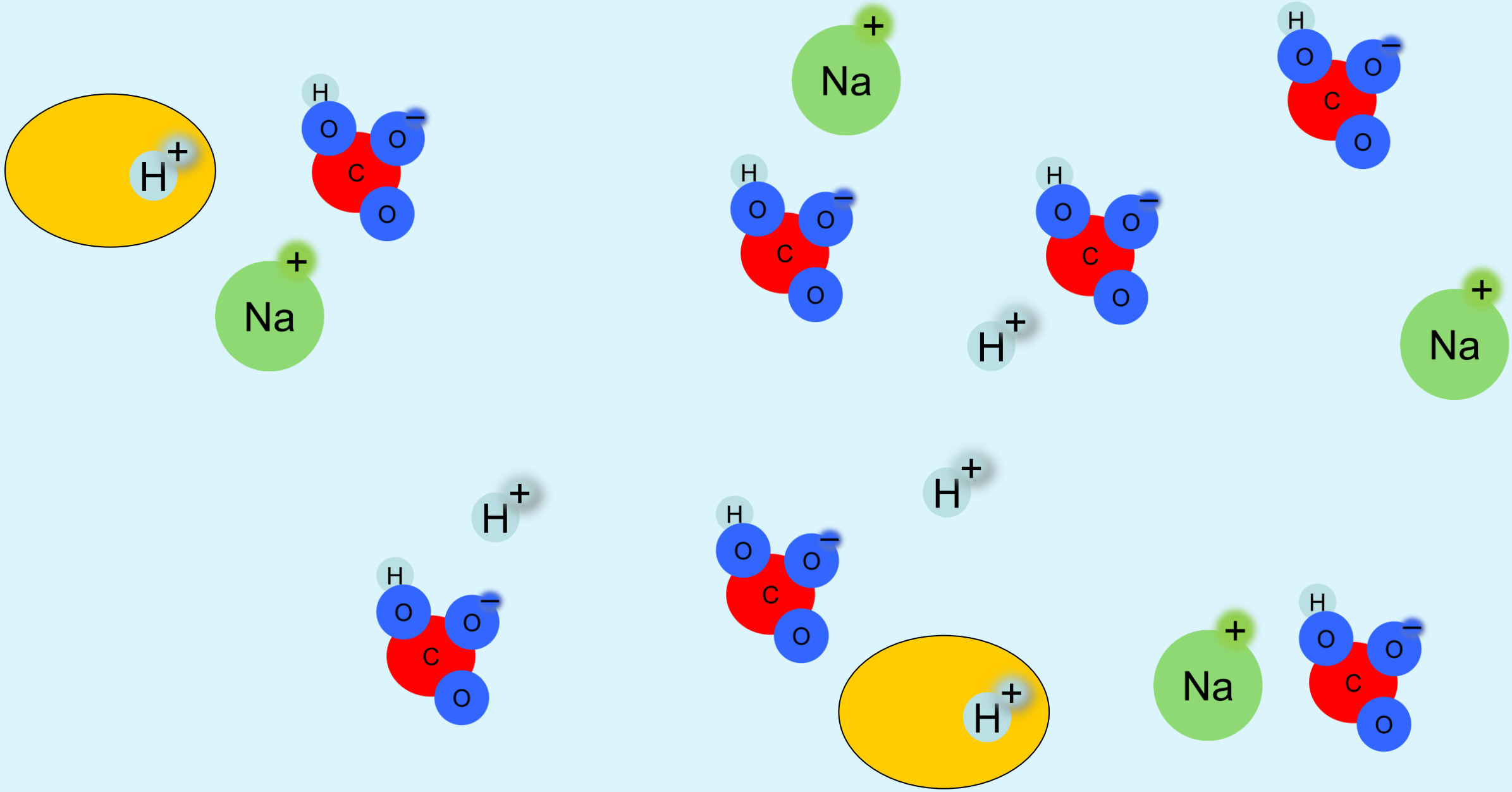


Bicarbonate
- base

What are buffers?

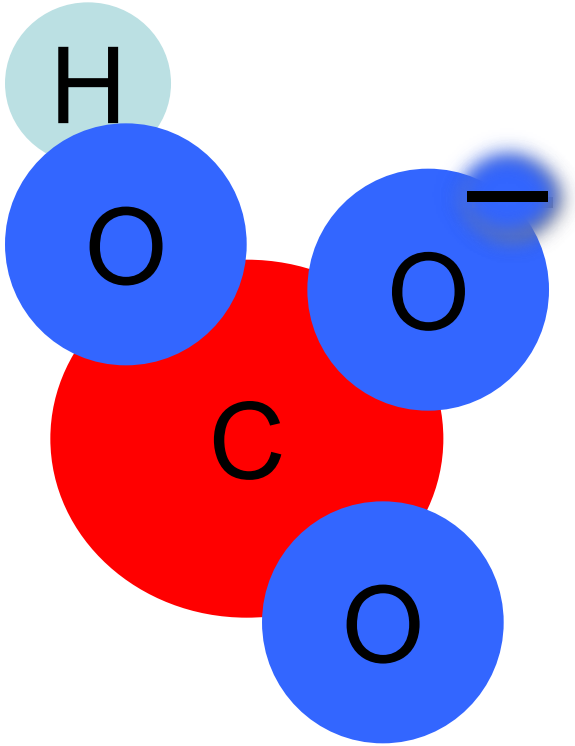
Buffers

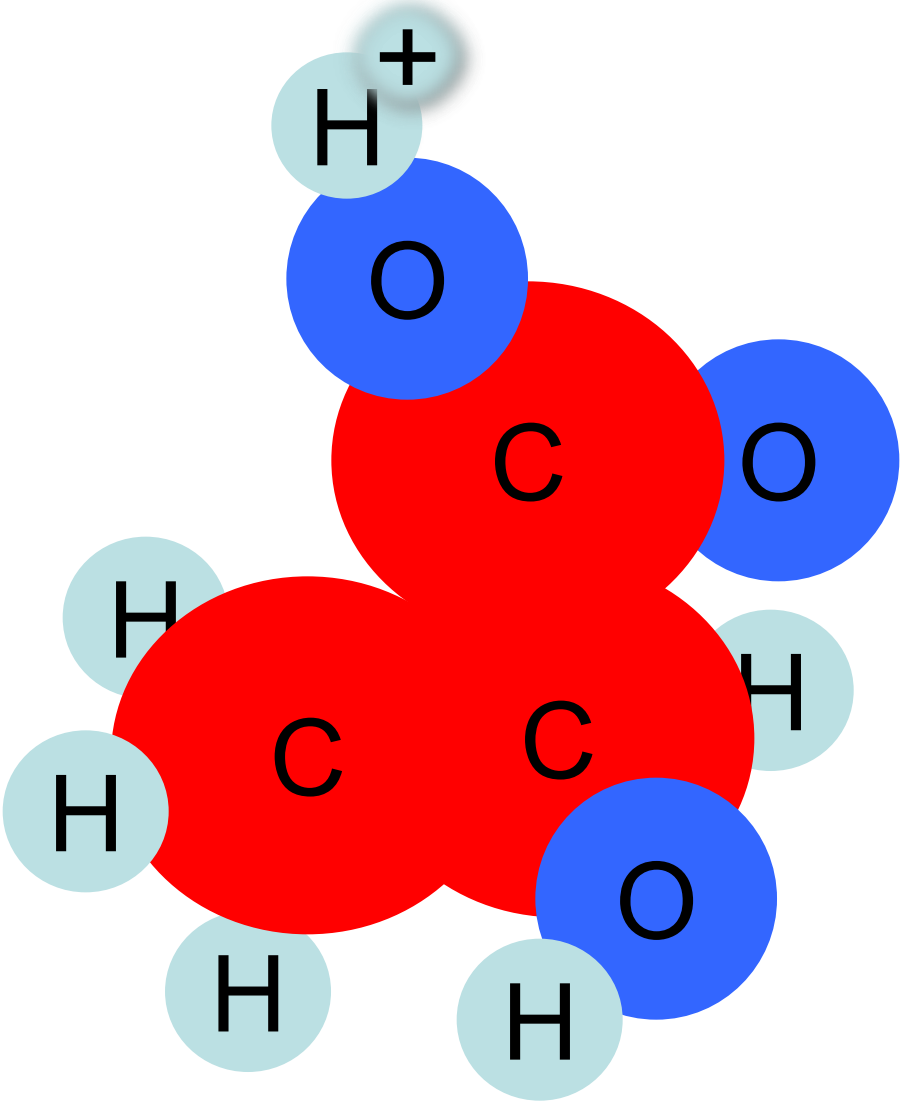
- Solution composed of a weak acid and its conjugate base (salt of the proper acid and a strong base eg. NaOH)
- Neutralize added H+ ions, thereby maintaining a stable pH.


H+
H
O
O
C
O
Na+
Na
H
O
O
C
O
H+
H
O
O
C
O
Na+
Na
H
O
O
C
O
H+
H
O
O
C
O
H+
H+
H
O
O
C
O
H+
Na+
Na
H
O
O
C
O

$NaHCO_3 \rightarrow Na^+ + HCO_3^-$

pH





THE THEORY OF NEUTRALITY REGULATION IN THE ANIMAL ORGANISM.

1908

By LAWRENCE J. HENDERSON.

[*From the Laboratory of Biological Chemistry of the Harvard Medical School.*]

WITH the results reported in the preceding paper, the work which has been carried on in this laboratory concerning equilibria of neutrality[1] during the last three years has progressed far enough to make possible a general discussion of the equilibria between phosphates and carbonates in the body of the warm-blooded animal. Such is the purpose of the present paper.

The Simple Equilibrium.

In the ionization reaction of a weak acid, $HA = \dot{H} + \bar{A}$, the conditions of equilibrium are defined by the equation of the concentration law,

$$k \cdot (HA) = (\dot{H}) \cdot (\bar{A}),$$

in which the enclosed quantities stand for concentrations, and k is the ionization constant of the acid. This equation may more conveniently be written as follows:

$$(\dot{H}) = k \times \frac{(HA)}{(\bar{A})}.$$

If an aqueous solution of such an acid and its sodium salt be prepared, it is evident that here the concentration of the undissociated molecules of the acid (HA) will be almost precisely equal to the total amount of acid present, for such an acid under these conditions will be hardly at all ionized. On the other hand, the concentration

[1] HENDERSON: This journal, 1906, xv, p. 257; 1908, xxi, p. 169; 1908, xxi, p. 173. HENDERSON and BLACK: This journal, 1907, xviii, p. 250; 1908, xxi, p. 420. FITZ, ALSBERG, and HENDERSON: This journal, 1907, xviii, p. 250. HENDERSON and WEBSTER: The journal of medical research, 1907, xvi, p. 1.

Henderson 1908

$$\underset{\text{(acid)}}{HA} \rightarrow H^+ + \underset{\text{(base)}}{HA^-}$$

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

Henderson – Hasselbalch

$$pH = 6{,}1 + log \frac{[HCO_3^-]}{PCO_2 \times 0{,}0307}$$

Constants: pK: -log[10] dissociation constant of carbonic acid

Quotient of CO_2 solubility

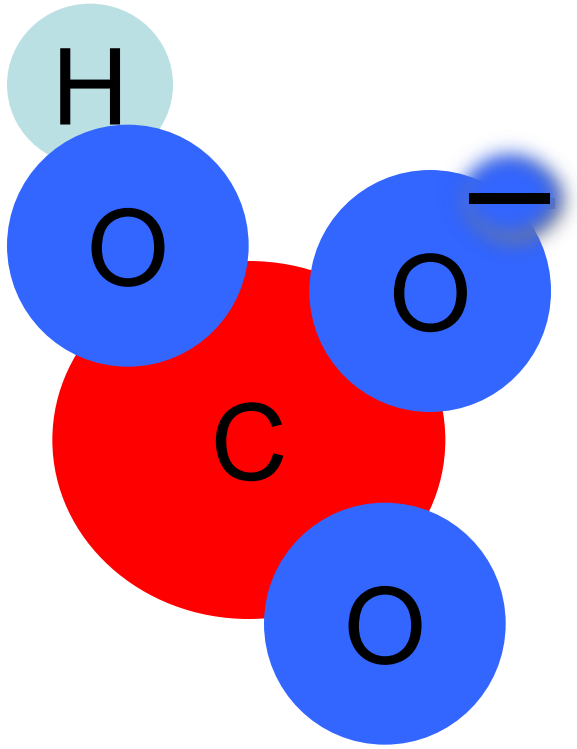
Henderson – Haselbalch

$$pH = pK \times log \frac{[HCO_3^-]}{P_aCO_2 \times 0{,}0307}$$

$$pH = 6{,}1 + log\,\frac{24}{1{,}2} = 7{,}4$$

- Adding 1,1 mmol/l of lactate:
- $[H^+]$ = 40 nmol/l ⟶ 1,14 mmol/l

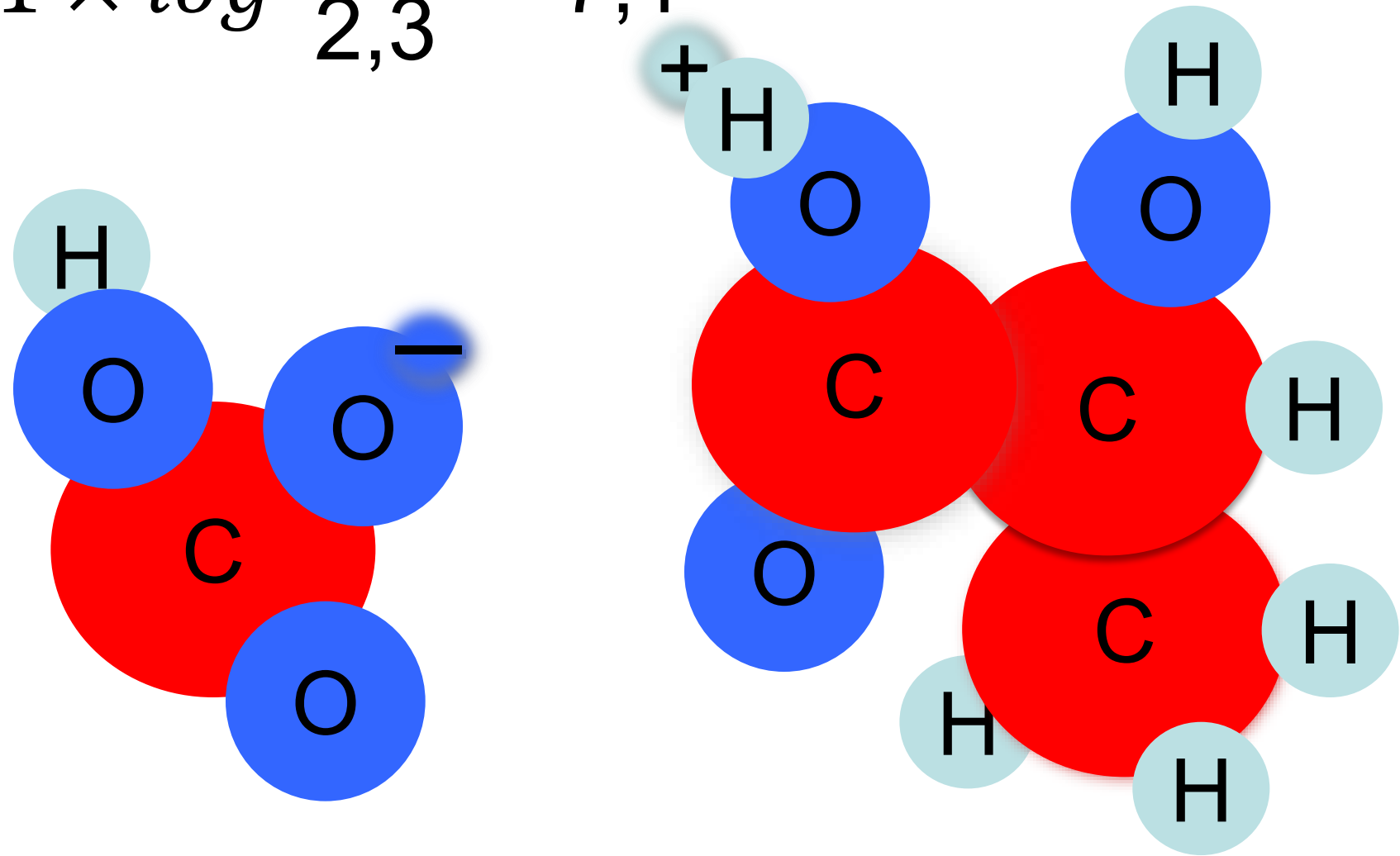
$$pH = -\log\,(0{,}00114) = 2{,}9$$

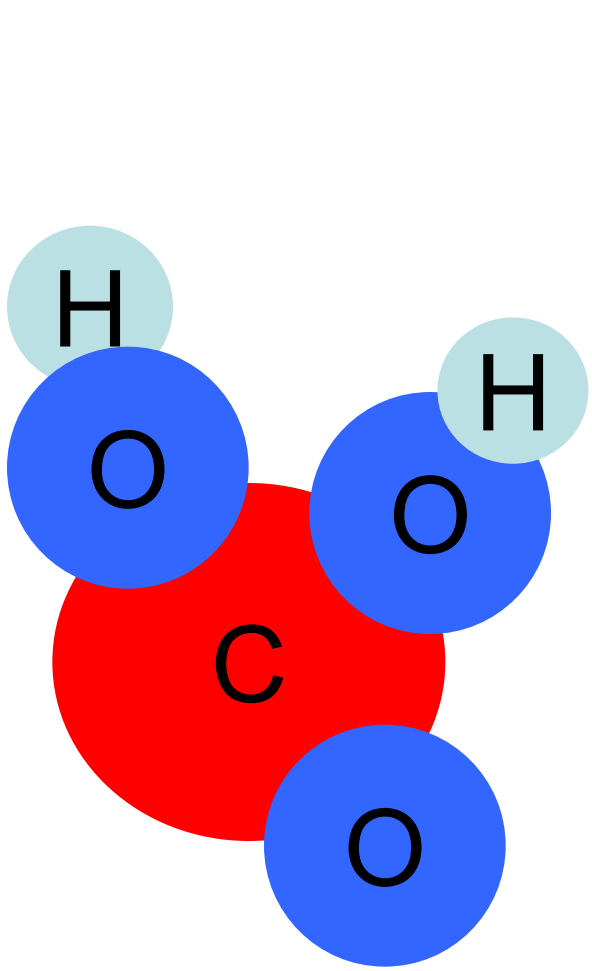

H
O
O
−
C
O

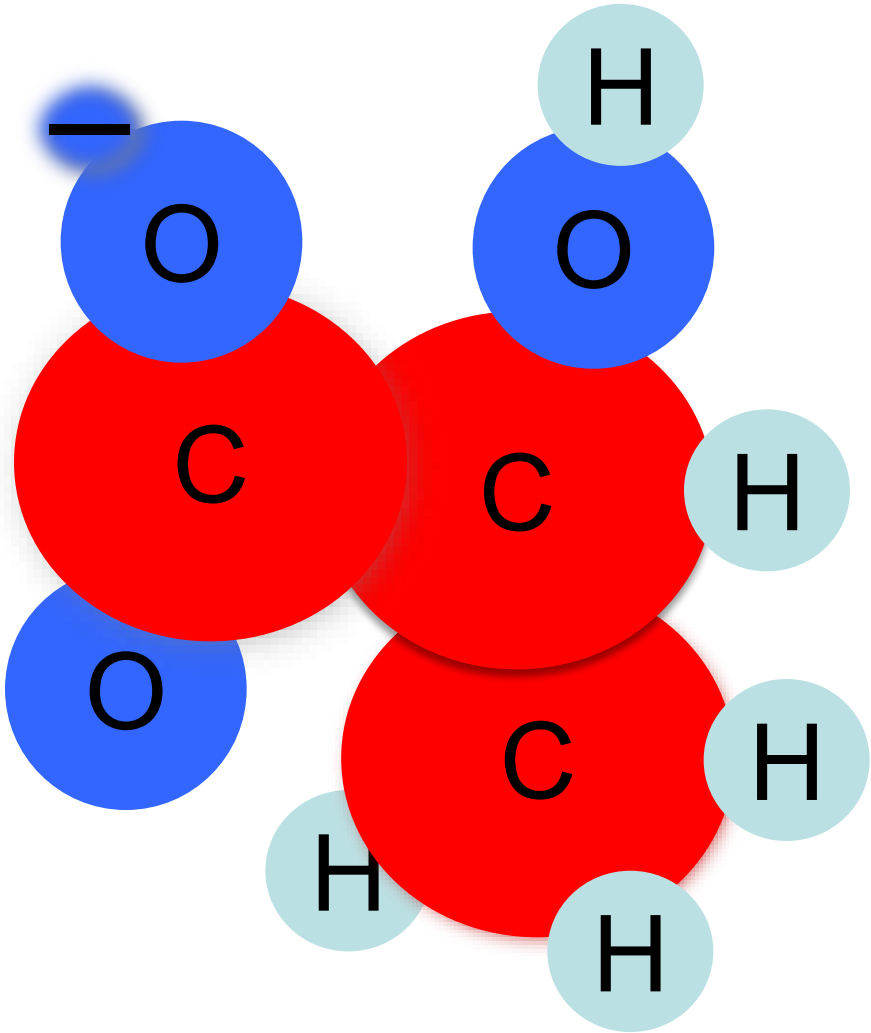
pH = -log (0,00114) = 2,9

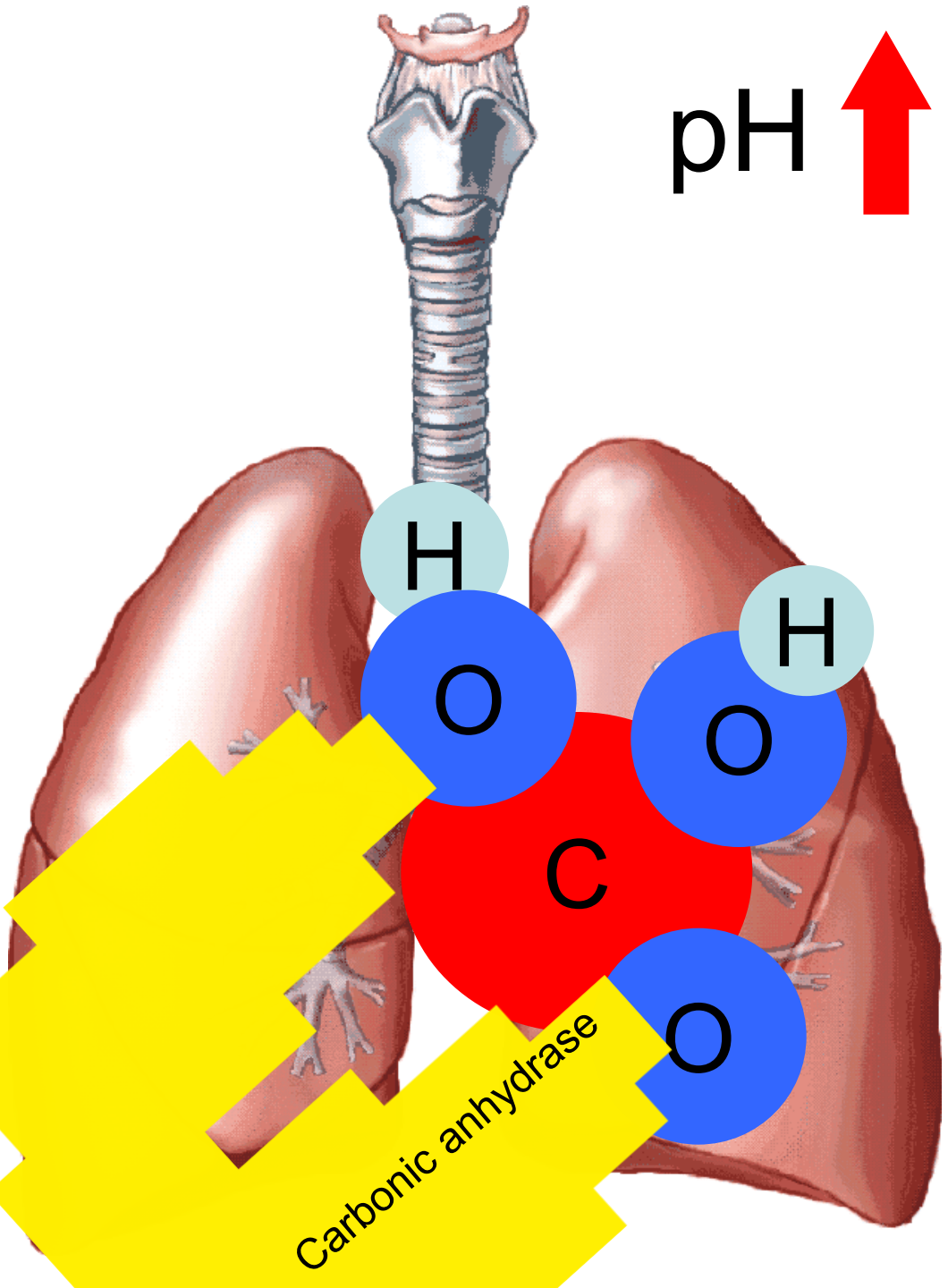
$pH = 6{,}1 \times log \frac{21{,}9}{2{,}3}$ = 7,1

pH




H
O
O
H
C
O


H
O
O
C
C
H
O
C
H
H
H



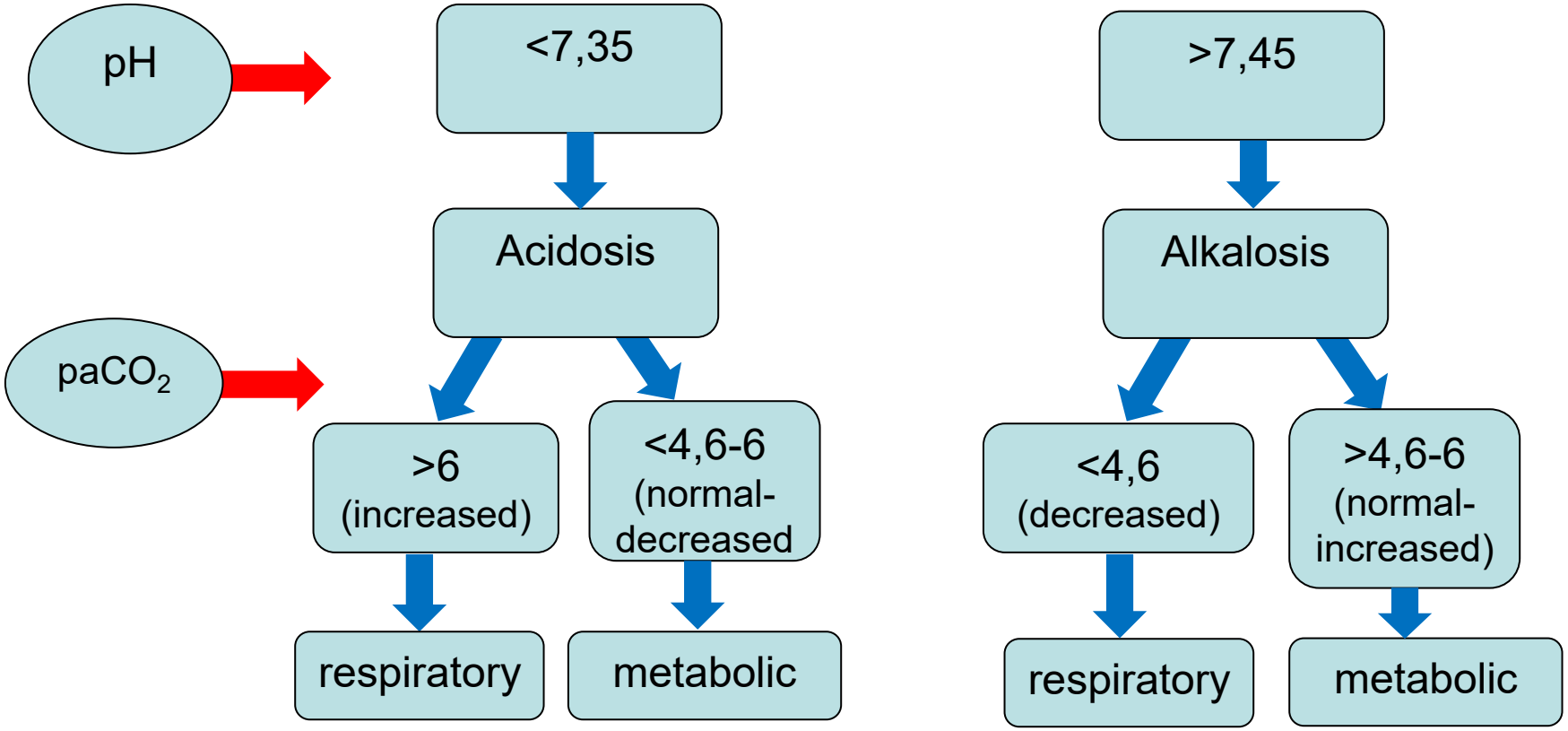
pH ↑

$$pH = 6{,}1 \times log \frac{21{,}9}{2{,}3} = 7{,}1$$

$$pH = 6{,}1 \times log \frac{21{,}9}{1{,}2} = 7{,}38$$

$$H_2CO_3 \longrightarrow CO_2 + H_2O$$

Classification of Acid-base disorders



Interpretation based on bicarbonate

Metabolic Acidosis ← $[HCO_3^-]$ 24 mmol/l → Metabolic Alkalosis

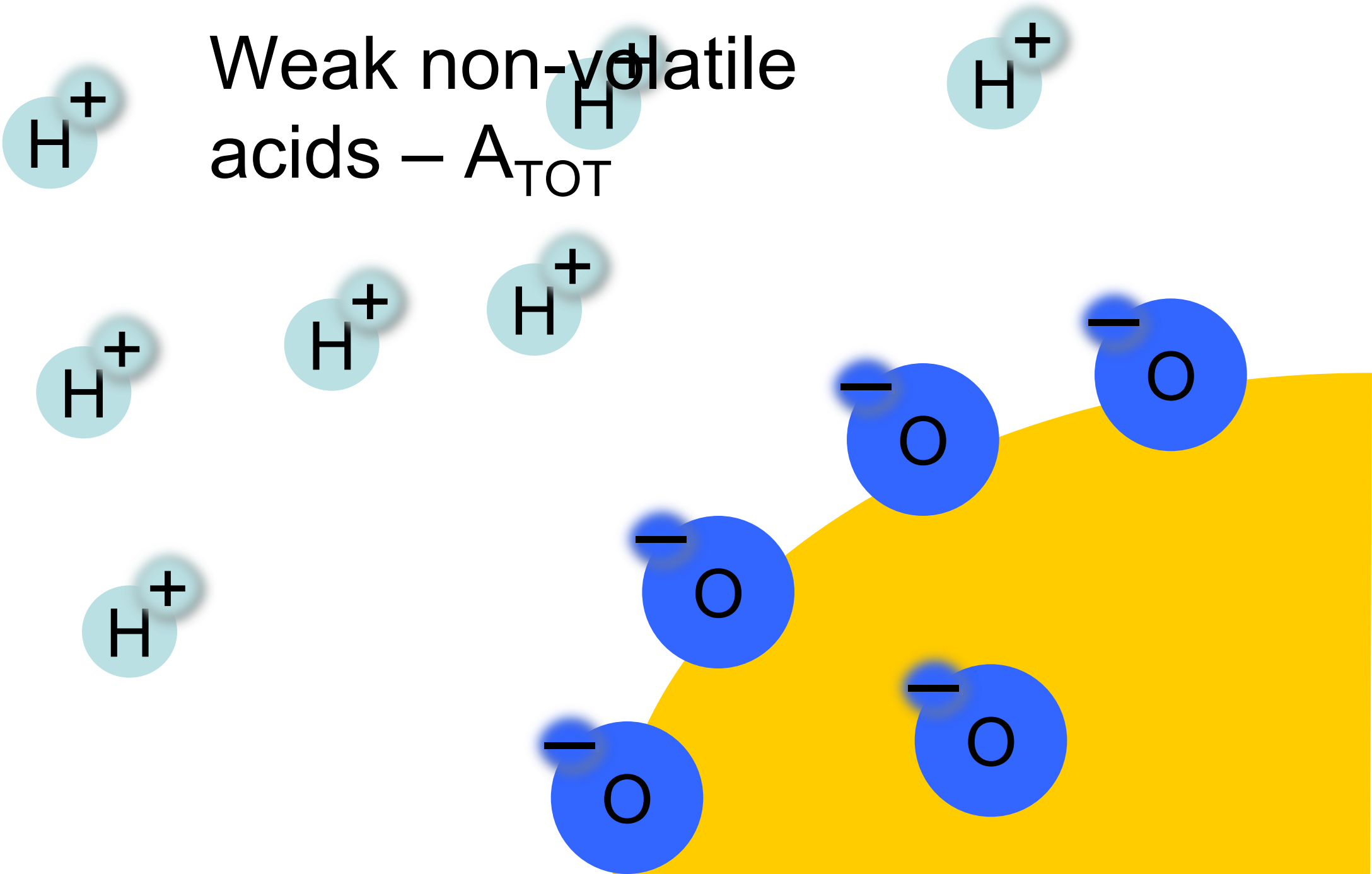
$CO_2 + H_2O \rightarrow H_2CO_3 \rightarrow H^+ + HCO_3^-$

$$[HCO_3^-] = K \frac{[H_2CO_3]}{[H^+]}$$

Buffer base

$$BB = [HCO_3^-] + A_{TOT}$$

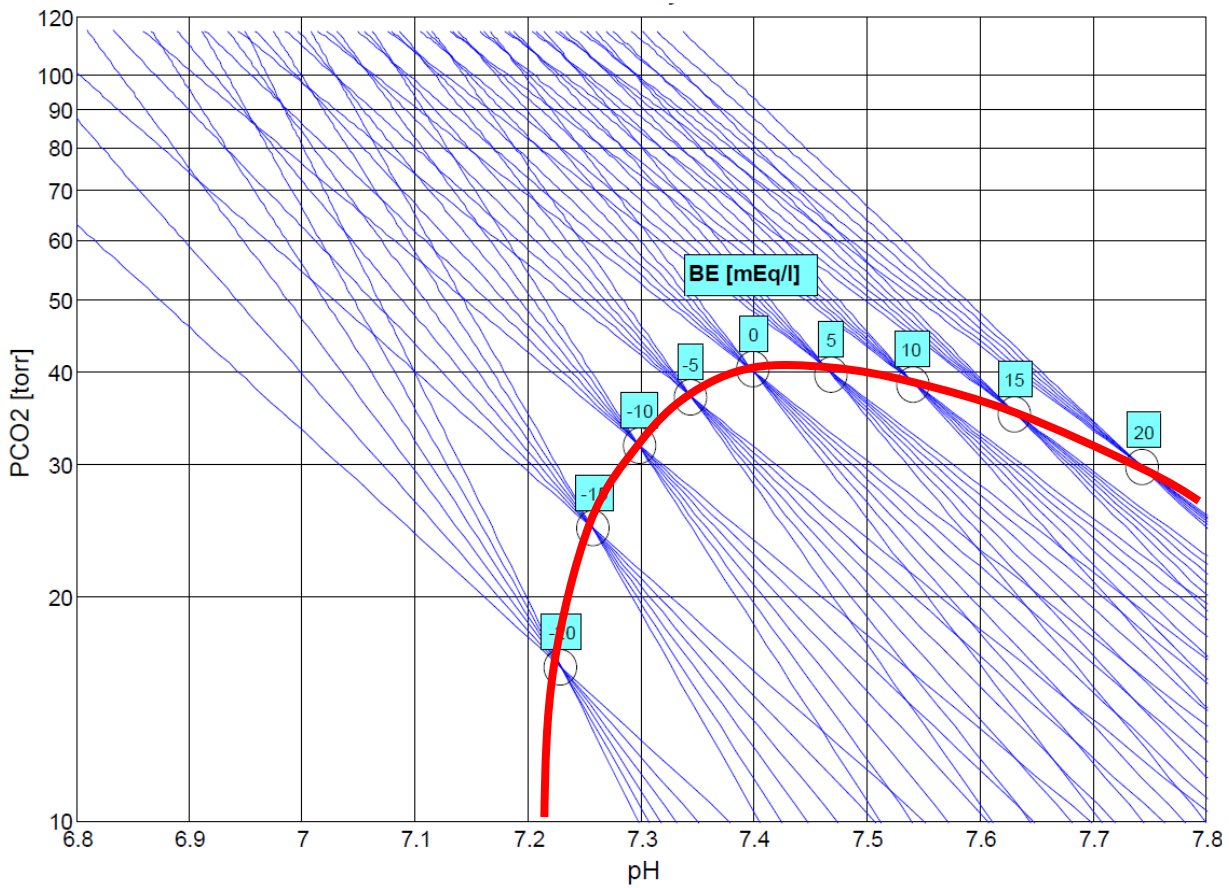
Weak non-volatile acids – A_{TOT}



Andersen, O. S., K. Engel, et al. (1960).

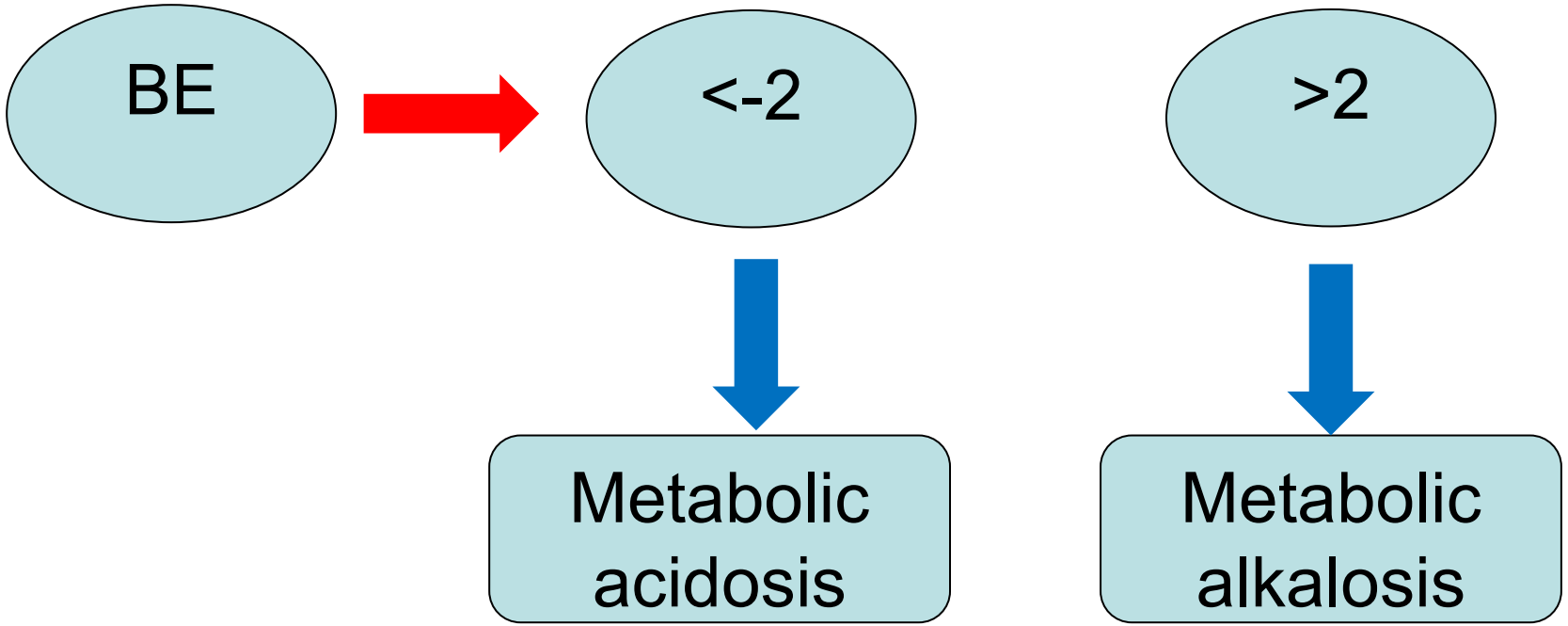
"A Micro method for determination of pH, carbon dioxide tension, base excess and standard bicarbonate in capillary blood."

Scand J Clin Lab Invest **12**: 172-6.

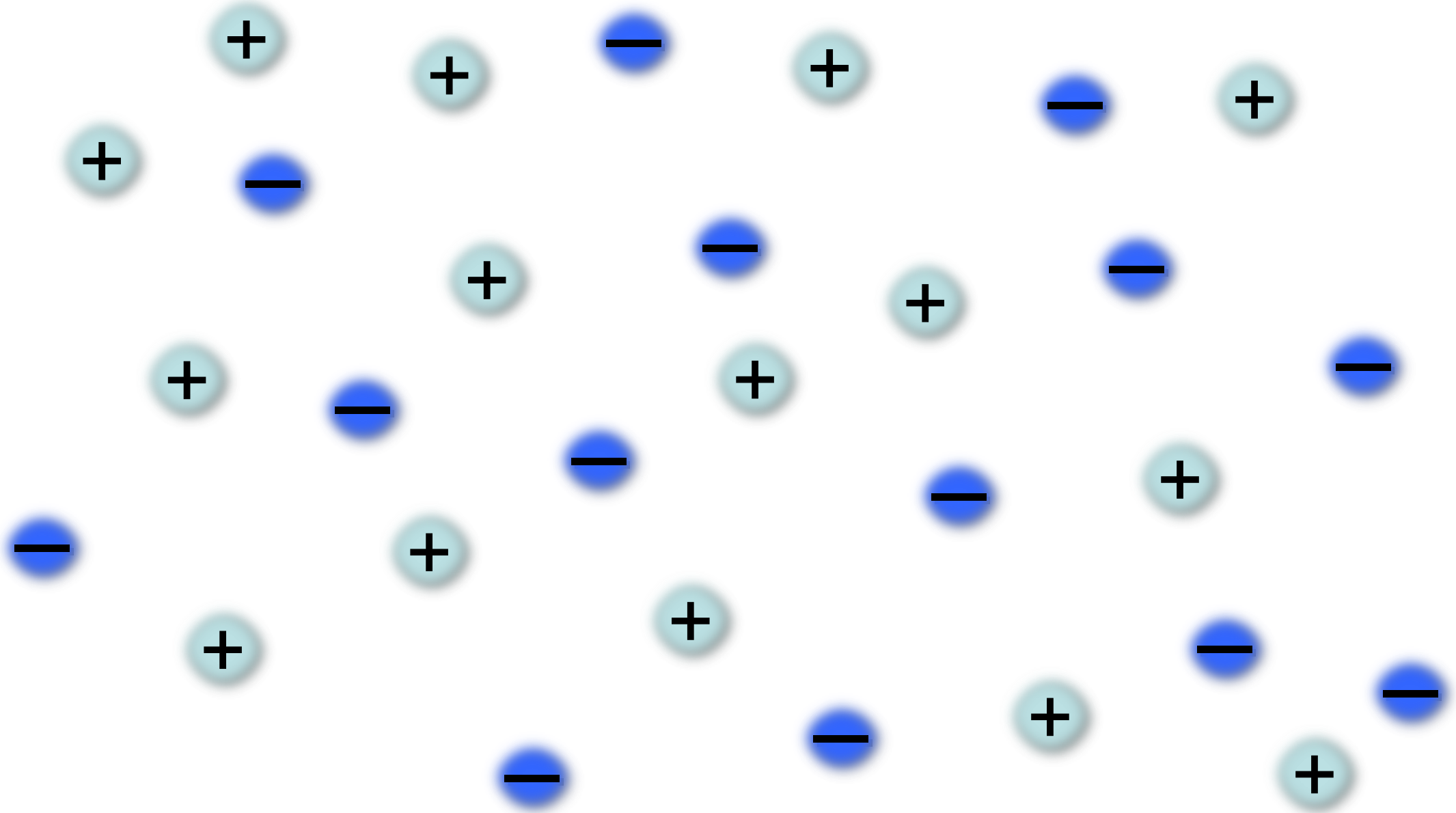


Base excess

- BE = BB (actual) – BB (normal)
- The amount of acid or base needed to normalize a blood sample's pH (7.40) under standard conditions (37°C and pCO_2 of 5 kPa).
- BE is negative, when there is a deficit of base in the sample (need to add)
- BE is positive, when there is an excess of base (need to add acid)
- Physiological value: -2 – +2 mmol/l

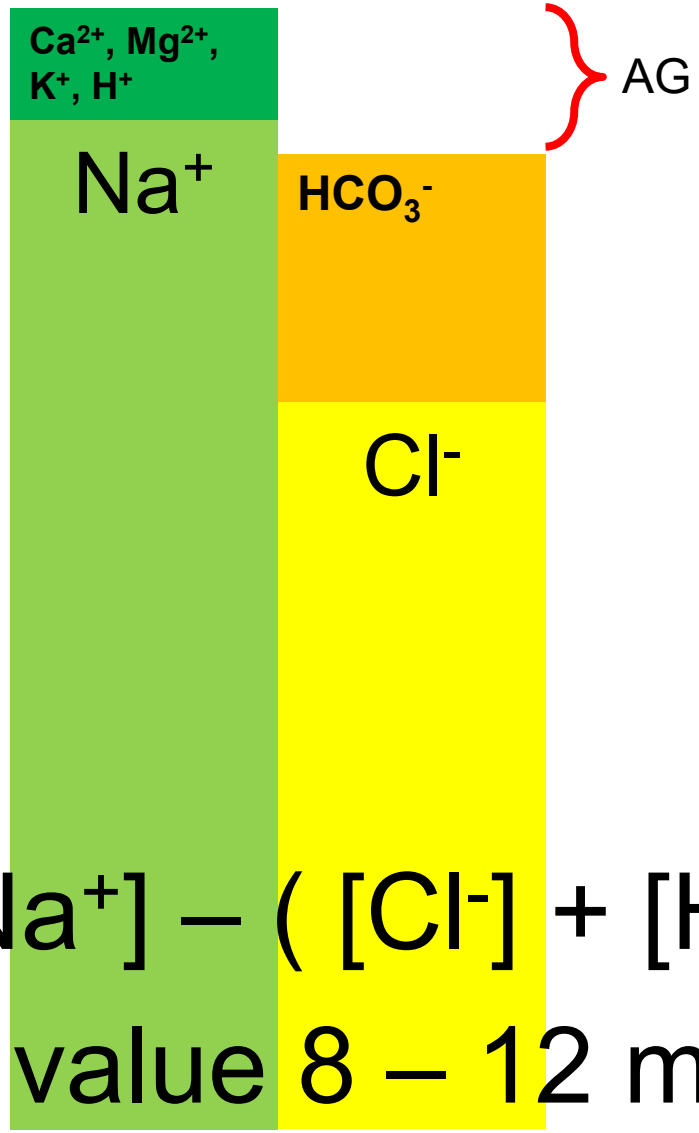

BE
<-2
>2
Metabolic acidosis
Metabolic alkalosis

Anion Gap

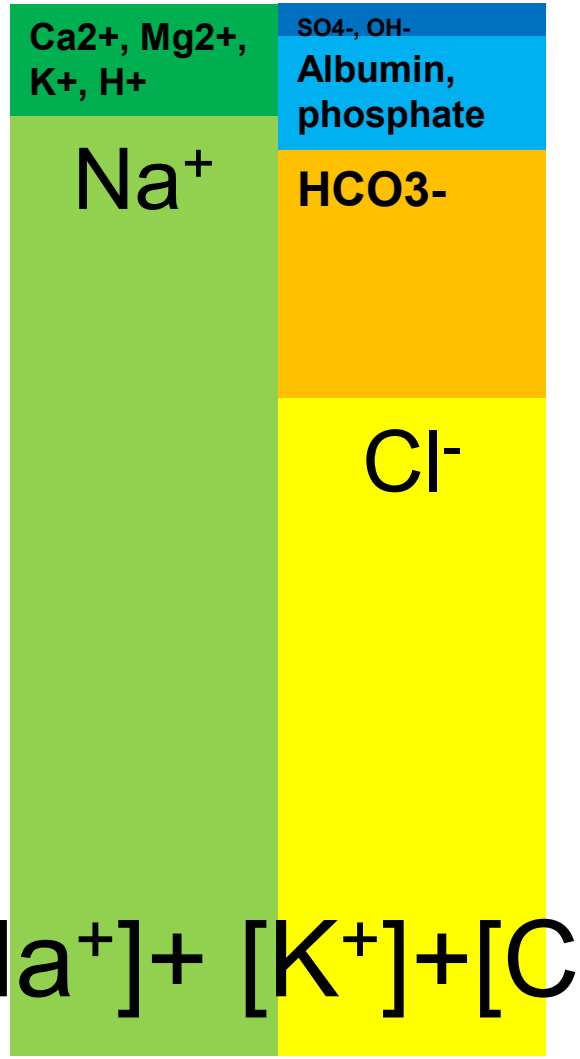


Principle of electroneutrality

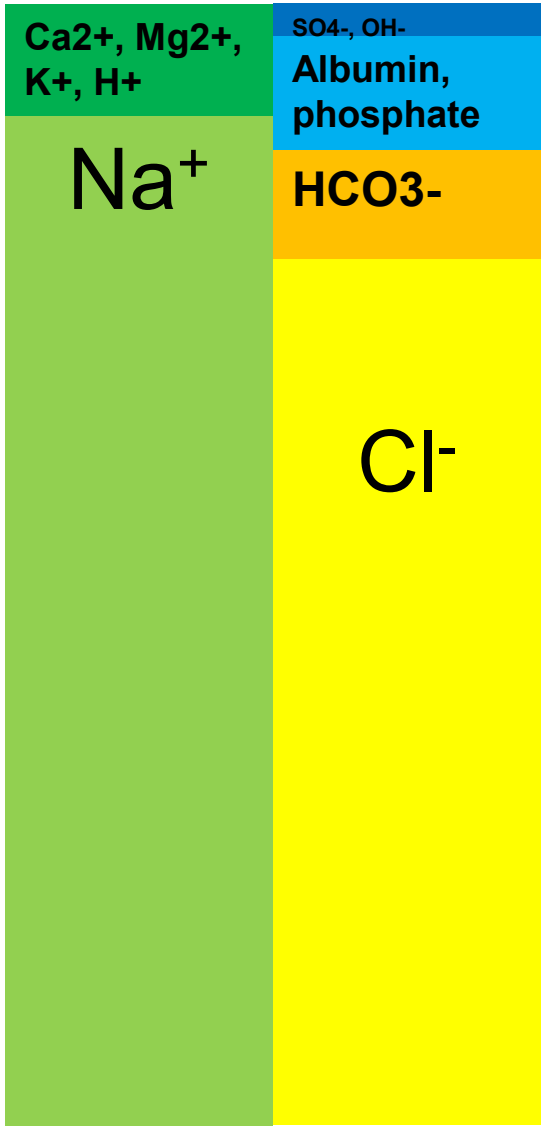
- In a given macroscopic system, the total amount of positive charge must equal the total amount of negative charge, ensuring a net electrical charge of zero

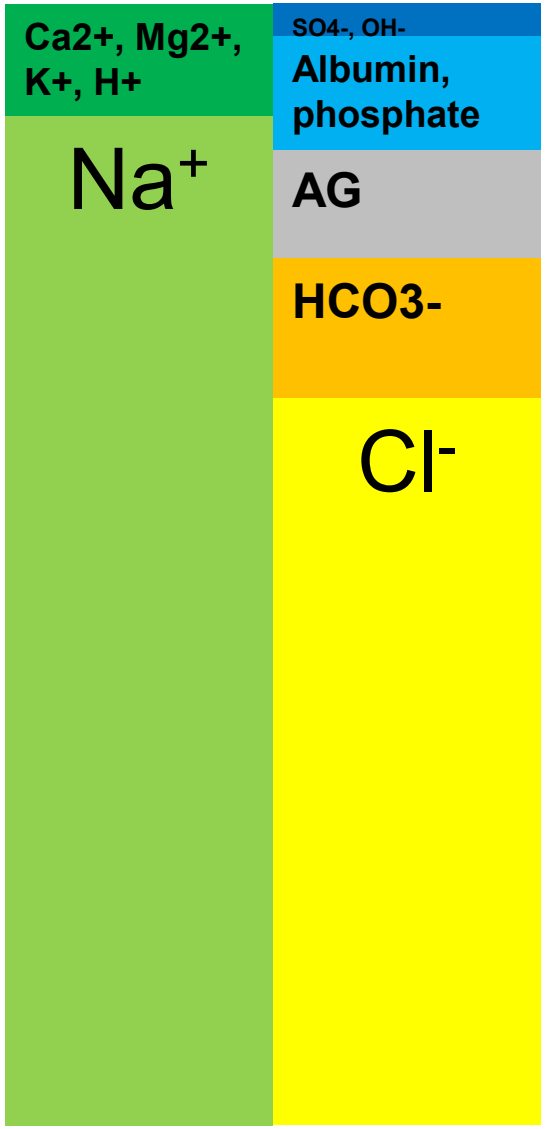


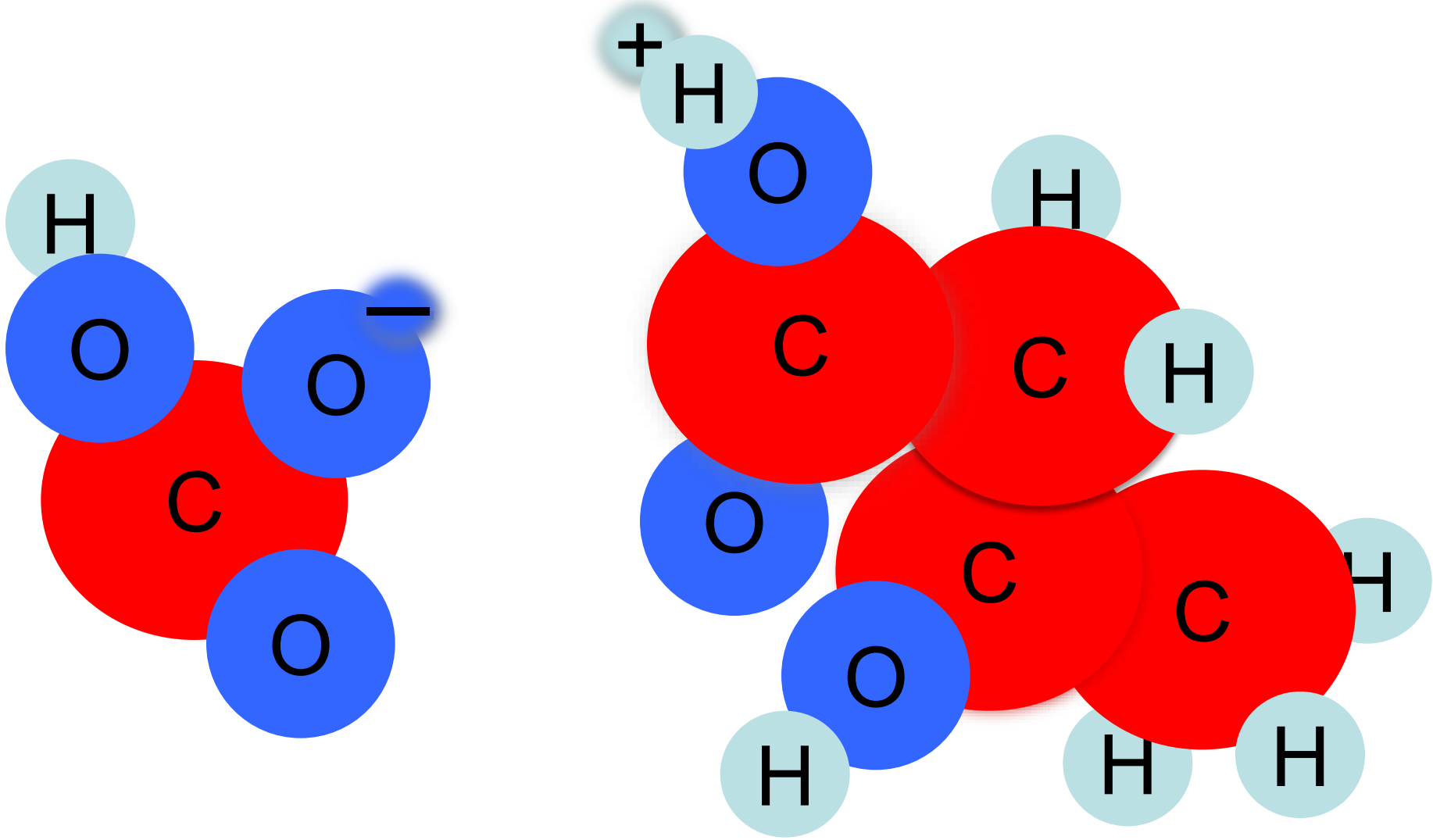
- AG = $[Na^+] – ([Cl^-] + [HCO_3^-])$
- Normal value 8 – 12 mmol/l



$$AG_c = ([Na^+]+ [K^+]+[Ca^{2+}]+ [Mg^{2+}])$$


Ca2+, Mg2+, K+, H+
Na+
SO4-, OH-
Albumin, phosphate
HCO3-
Cl-


Ca2+, Mg2+, K+, H+
Na+
SO4-, OH-
Albumin, phosphate
AG
HCO3-
Cl-


H
O
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H
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C
H
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H
H
H

Metabolic acidsis

High AG metabolic acidosis - HAGMA

- M — Methanol
- U — Uremia
- D — Diabetic/alcoholic ketoacidosis
- P — Propylene glycol
- I — Iron, isoniazid
- L — Lactate
- E — Ethylene glycol/ethanol
- S — salicylates

Normal AG metabolic acidosis – hyperchloremic

1. GIT secretion
2. RTA
3. Iatrogenic

Conclusion

- Interpretation based on pH and pCO_2
- Bicarbonate is a powerful buffer
- Metabolic disorders are quantified either by the concentration of bicarbonate or by BE
- Unmeasured anions are evaluated by the AG

Thank You