Arrhenius Acids and Bases

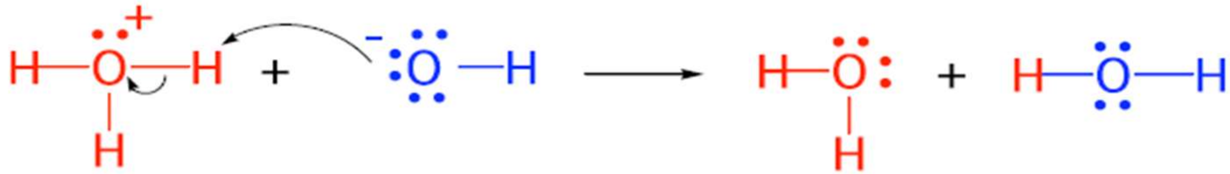
acid: substance that produces H_3O^+ ions aqueous solution
base: substance that produces OH^- ions in aqueous solution

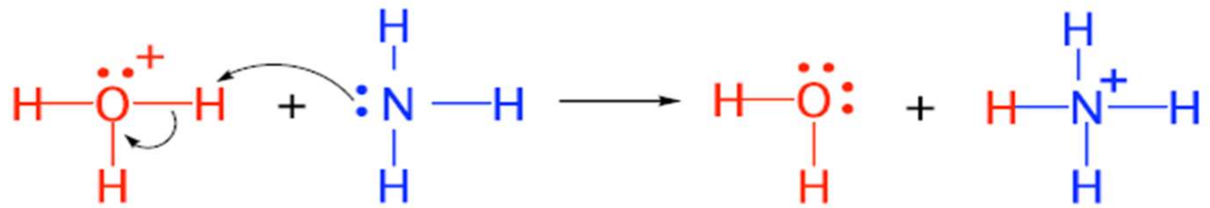
$$H^+(aq) + H_2O(l) \longrightarrow H_3O^+(aq)$$

$$KOH \xrightarrow{H_2O} K^+ + {}^-OH(aq) \text{ ion}$$

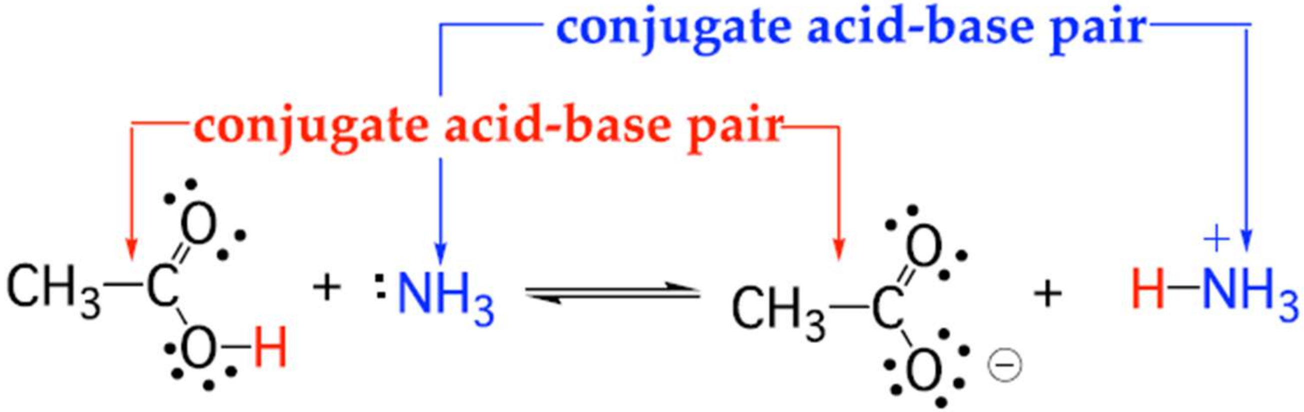
Brønsted-Lowry Definitions

Acid: a proton donor
Base: a proton acceptor



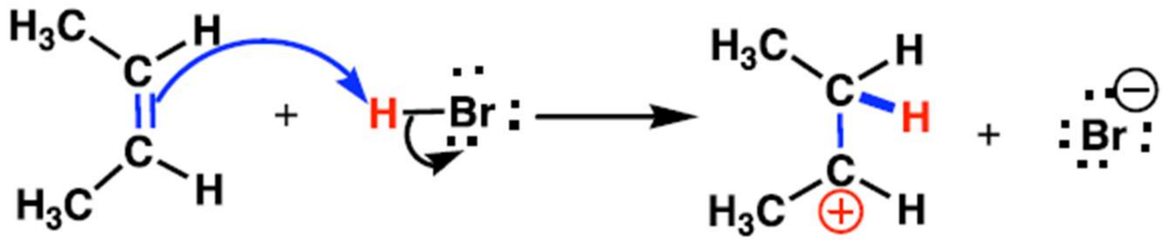


Conjugate Acids & Bases **Brønsted-Lowry** does not require water only **H+** transfer



Pi Electrons As Basic Sites

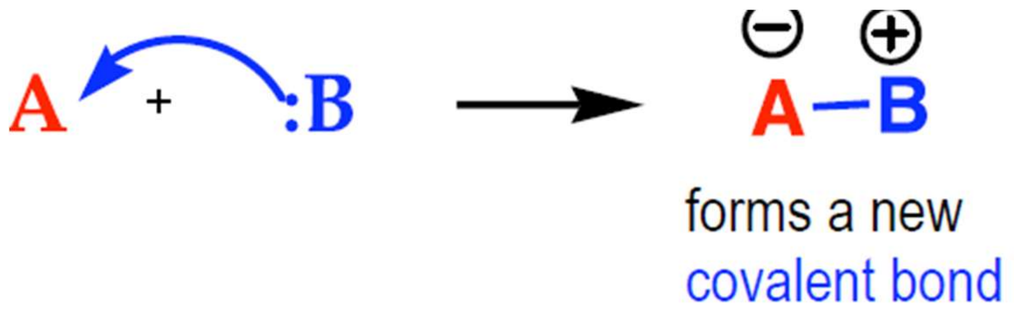
pi electrons - base - donate a pair of e-



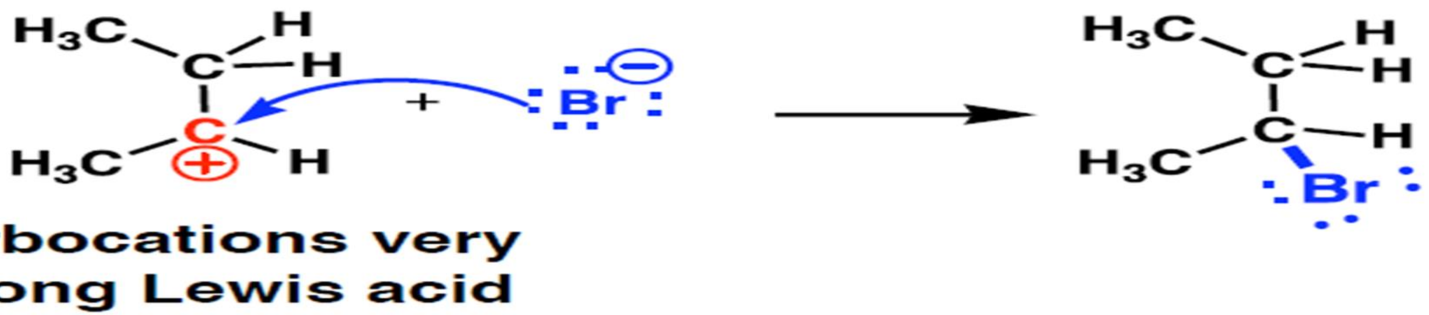
Result: formation of a carbocation,
C has 6-e⁻s, ∴ +1charge

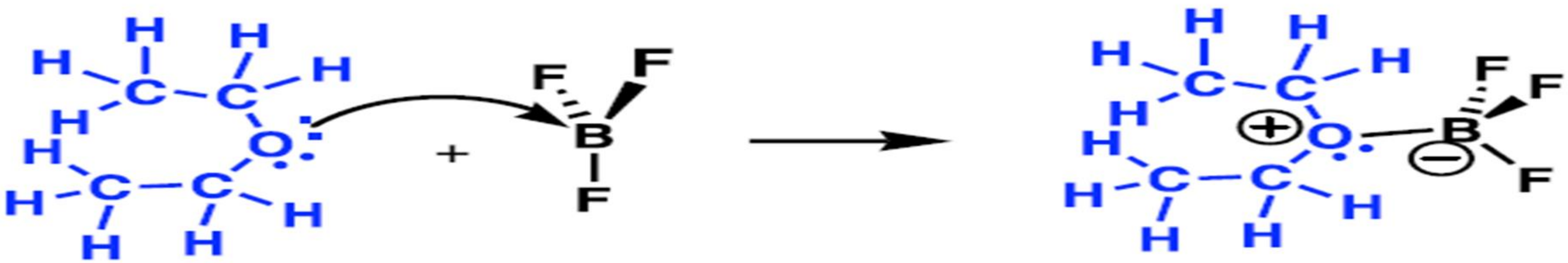
Lewis Acids and Bases

Lewis acid: molecule/ion that **accepts a pair of electrons**
Lewis base: molecule/ion that **donates a pair of electrons**



examples



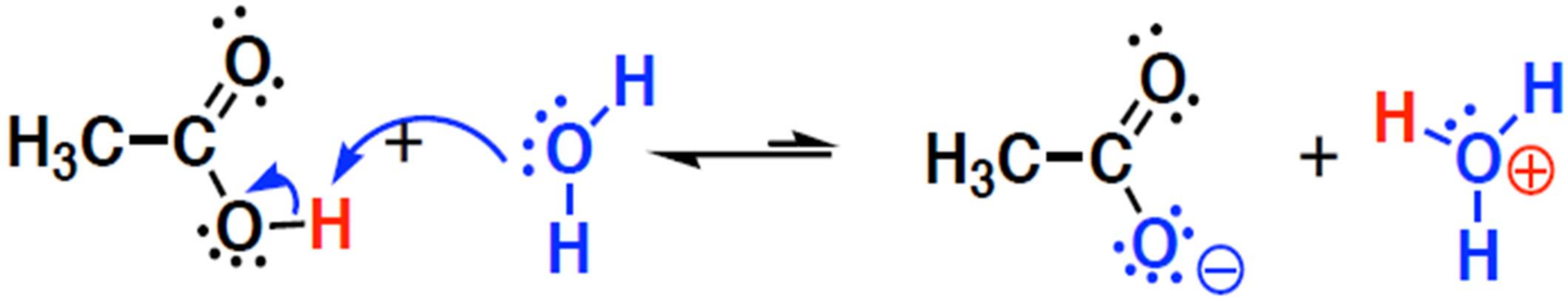


Acids & Base Strengths

How is acid [base] strength expressed/compared?

By equilibrium constant

e.g.: dissociation (ionization) of acetic acid



$$K_{eq} = \frac{[H_3O^+][A^-]}{[HA][H_2O]}$$

$$K_{\mathbf{a}} = K_{eq}[H_2O] = \frac{[H_3O^+][A^-]}{[HA]}$$

$$pKa = -\log Ka$$

	Acid	Formula	pK_a	Conjugate Base	
Weaker acid	Ethane	CH_3CH_3	51	$CH_3CH_2^-$	Stronger conjugate base
	Ethylene	$CH_2=CH_2$	44	$CH_2=CH^-$	
	Ammonia	NH_3	38	NH_2^-	
	Hydrogen	H_2	35	H^-	
	Acetylene	$HC\equiv CH$	25	$HC\equiv C^-$	
	Ethanol	CH_3CH_2OH	15.9	$CH_3CH_2O^-$	
	Water	H_2O	15.7	HO^-	
	Methylammonium ion	$CH_3NH_3^+$	10.64	CH_3NH_2	
	Bicarbonate ion	HCO_3^-	10.33	CO_3^{2-}	
	Phenol	C_6H_5OH	9.95	$C_6H_5O^-$	
	Ammonium ion	NH_4^+	9.24	NH_3	
	Hydrogen sulfide	H_2S	7.04	HS^-	
	Carbonic acid	H_2CO_3	6.36	HCO_3^-	
	Acetic acid	CH_3COOH	4.76	CH_3COO^-	
	Benzoic acid	C_6H_5COOH	4.19	$C_6H_5COO^-$	3.18 HF
	Phosphoric acid	H_3PO_4	2.1	$H_2PO_4^-$	
	Hydronium ion	H_3O^+	-1.74	H_2O	
	Sulfuric acid	H_2SO_4	-5.2	HSO_4^-	
	Hydrogen chloride	HCl	-7	Cl^-	
	Hydrogen bromide	HBr	-8	Br^-	
Stronger acid	Hydrogen iodide	HI	-9	I^-	Weaker conjugate base

Buffer system

buffer system is a solution that resists a change in pH when acids or bases are added.

How does a buffer work?

When a strong base is added, the acid present in the buffer neutralizes the hydroxide ions

When a strong acid is added, the base present in the buffer neutralizes the hydronium ions

Buffer systems in the body

Carbonic acid bicarbonate buffer system

Phosphate buffer system

Protein buffer system.

1-Bicarbonate buffer.

Bicarbonate buffer system The pH of a solution ultimately depends on the concentration of H+ ions in the solution. Changing the concentration, by the addition or removal of H+ ions, can impact the pH of a solution. The bicarbonate buffering system is an crucial buffer system in the acid-base homeostasis of all living things. The main role of the bicarbonate system is to regulate and control the pH of blood and counteract any force that will alter the pH. A buffer system exists to help neutralize the blood if excess hydrogen or hydroxide ions are produced. An acid-base buffer consists of a weak acid

(H_2CO), CO_2 (g) + H_2O (l) ↔ H_2CO_3 (aq)↔ HCO_3^-(aq) + H^+(aq)

2- **Protein Buffer System**

- Plasma and intracellular proteins are the body’s most plentiful and powerful buffers
- Some amino acids of proteins have:
- Free organic acid groups (weak acids)
- Groups that act as weak bases (e.g., amino groups)
- Amphoteric molecules are protein molecules that can function as both a weak acid and a weak base

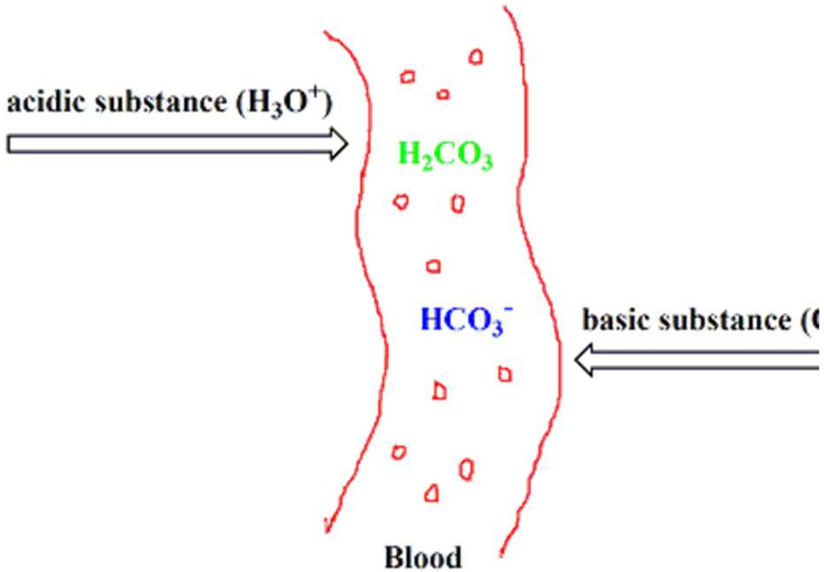
3- **Phosphate Buffer System**

- Nearly identical to the bicarbonate system
- Its components are:
- Sodium salts of dihydrogen phosphate ($H_2PO_4^-$), a weak acid
- Monohydrogen phosphate (HPO_4^-), a weak base
- This system is an effective buffer in urine and intracellular fluid

Buffer systems of blood

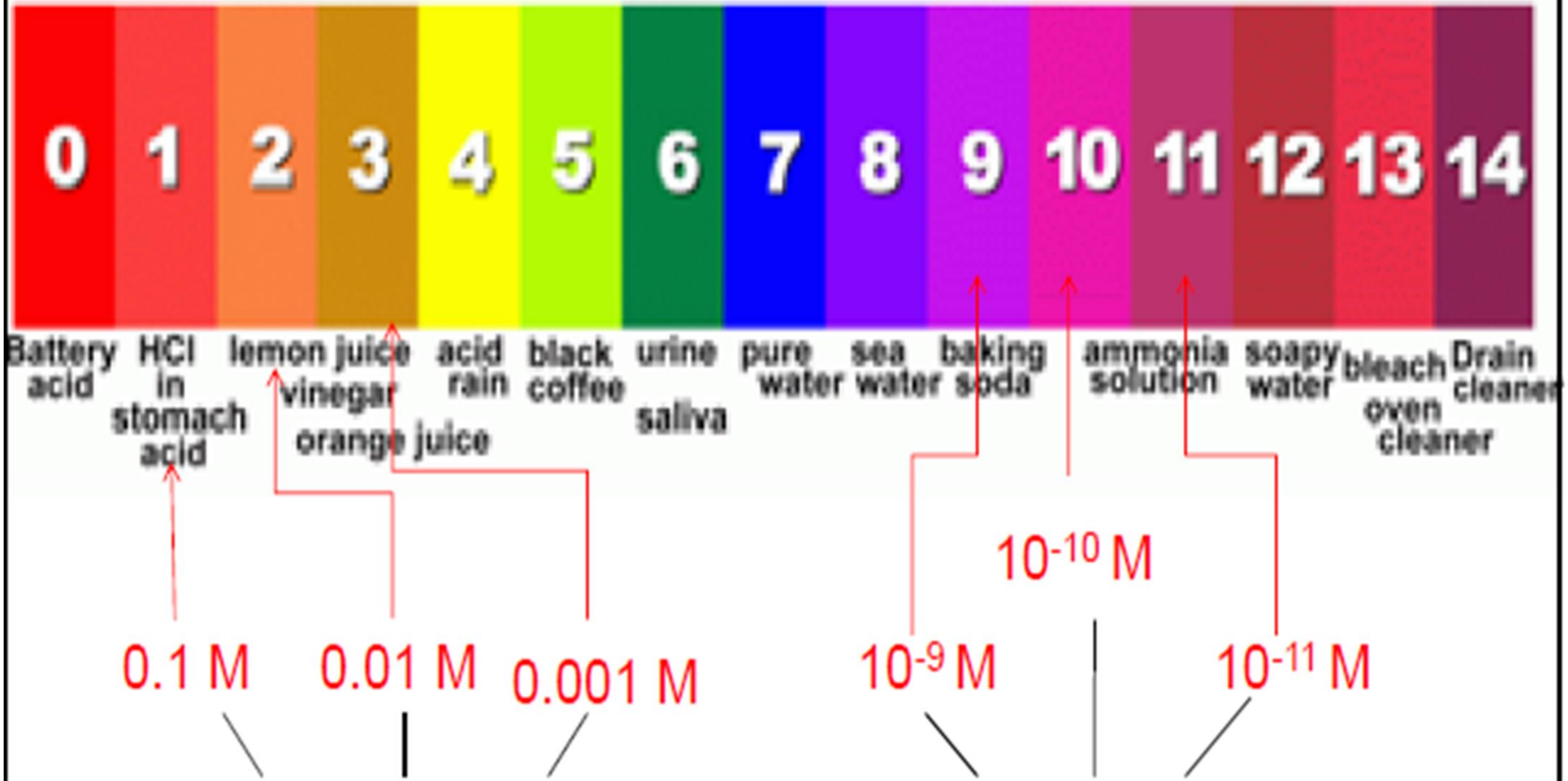
Maintaining a constant blood pH is critical for the proper functioning of our body. The buffer that maintains the pH of human blood involves a carbonic acid

$$\underset{\text{weak acid}}{H_2CO_3} + H_2O \rightleftharpoons \underset{\text{conjugate base}}{HCO_3^-} + H_3O^+$$



Recall the pH scale

$$pH = -\log [H^+]$$

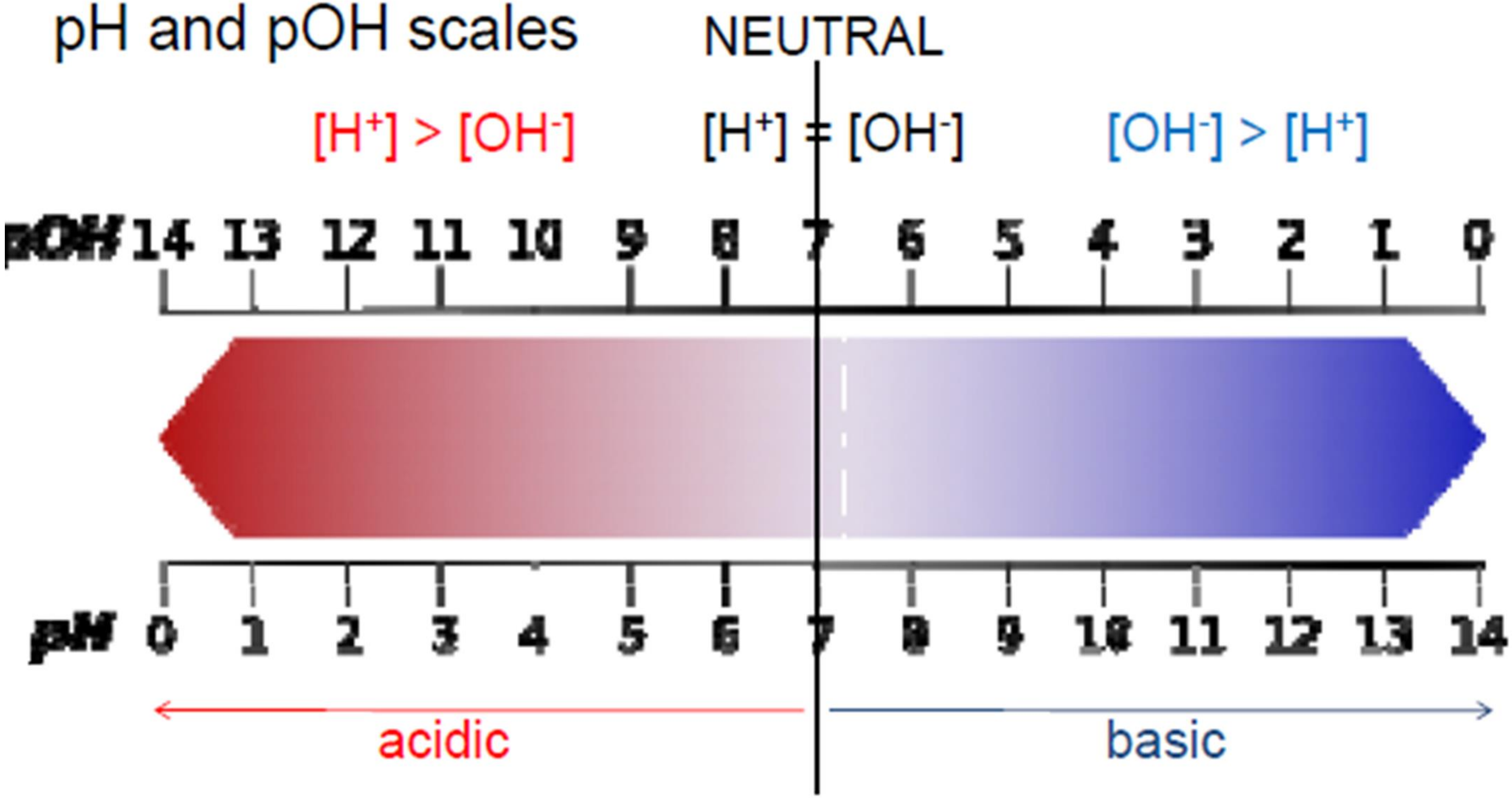


For an increase in 1 pH unit, the solution's $[H^+]$ decreases 10x.

Acids, Bases and Salts

Hebden – Unit 4 (page 109-182)

pH and pOH scales



Convert between pH and pOH: $pH + pOH = 14$ at 25°C

Calculate the pH of 0.30 M Na_2CO_3.

Step 1: Identify the ions involved.

$$Na_2CO_3\ (s) \rightarrow 2\ Na^+\ (aq) + CO_3^{2-}\ (aq)$$

Spectator ion

Found on the right side of the acid table; undergo base hydrolysis

Calculate the pH of 0.30 M Na_2CO_3.

Step 2: Write the net ionic equation for the hydrolysis. Set up ICE table.

$$CO_3^{2-}(aq) + H_2O(\ell) \rightleftharpoons HCO_3^-(aq) + OH^-(aq)$$

	$CO_3^{2-}(aq)$ +	$H_2O(\ell)$ $\rightleftharpoons$	$HCO_3^-(aq)$ +	$OH^-(aq)$
[I]	0.30		0	0
[C]	- x		+ x	+ x
[E]	0.30 - x		x	x

Stoichio-metric ratio

Calculate the pH of 0.30 M Na_2CO_3.

Step 3: Determine the K_b for the base hydrolysis.

$$CO_3^{2-}(aq) + H_2O(\ell) \rightleftharpoons HCO_3^-(aq) + OH^-(aq)$$

Look up in the acid table the K_a for HCO_3^-

$HCO_3^- \rightleftharpoons H^+ + CO_3^{2-}$ .... 5.6×10^{-11}

$$K_b = \frac{K_w}{K_a(HCO_3^-)} = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-11}} = 1.786 \times 10^{-4}$$

Calculate the pH of 0.30 M Na_2CO_3.

Step 4: Write the K_b expression for the base hydrolysis.

$$CO_3^{2-}(aq) + H_2O\ (\ell) \rightleftharpoons HCO_3^-\ (aq) + OH^-\ (aq)$$

$$K_b = \frac{[HCO_3^-]\cdot[OH^-]}{[CO_3^{2-}]}$$

Calculate the pH of 0.30 M Na_2CO_3.

Step 5: Substitute in the equilibrium concentrations from the ICE table.

	$CO_3^{2-}(aq)$	$+ H_2O\,(\ell)$ $\rightleftharpoons$	$HCO_3^-\,(aq)$	$+ OH^-\,(aq)$
[I]	0.30		0	0
[C]	- x		+ x	+ x
[E]	0.30 - x		x	x

$$K_b = \frac{[HCO_3^{-}]\cdot[OH^{-}]}{[CO_3^{2-}]} = \frac{x^2}{0.30 - x} = 1.786 \times 10^{-4}$$

Calculate the pH of 0.30 M Na_2CO_3.

Step 6: Solve for x, where x = $[OH^-]$.

$$K_b = \frac{[HCO_3^-]\cdot[OH^-]}{[CO_3^{2-}]} = \frac{x^2}{0.30 - x} = 1.786 \times 10^{-4}$$

$$\frac{x^2}{0.30} = 1.786 \times 10^{-4}$$

Use assumption $0.30 - x \cong 0.30$
because
$[0.30] > 1000 \times 1.786 \times 10^{-4}$

$[OH^-] = x = 7.319 \times 10^{-3}$ M

Calculate the pH of 0.30 M Na_2CO_3.

Step 7: From the [OH^-], calculate the pOH.

$$pOH = -\log (7.319 \times 10^{-3}) = 2.1355$$

Step 8: Convert pOH to pH.

$$pH = 14 – pOH = 14 – 2.1355 = 11.86$$

Step 9: Check that your answer makes sense.

We expect CO_3^{2-} ions will undergo base hydrolysis. pH > 7.

Home work

Dissolve 45.0 g of NH_4Cl in 1.50 L of water. What is the pH of this solution? Molar mass of NH_4Cl is 53.5 g/mole.

Buffer disorders of acid-base balance can be distinguished

Respiratory Acidosis and Alkalosis

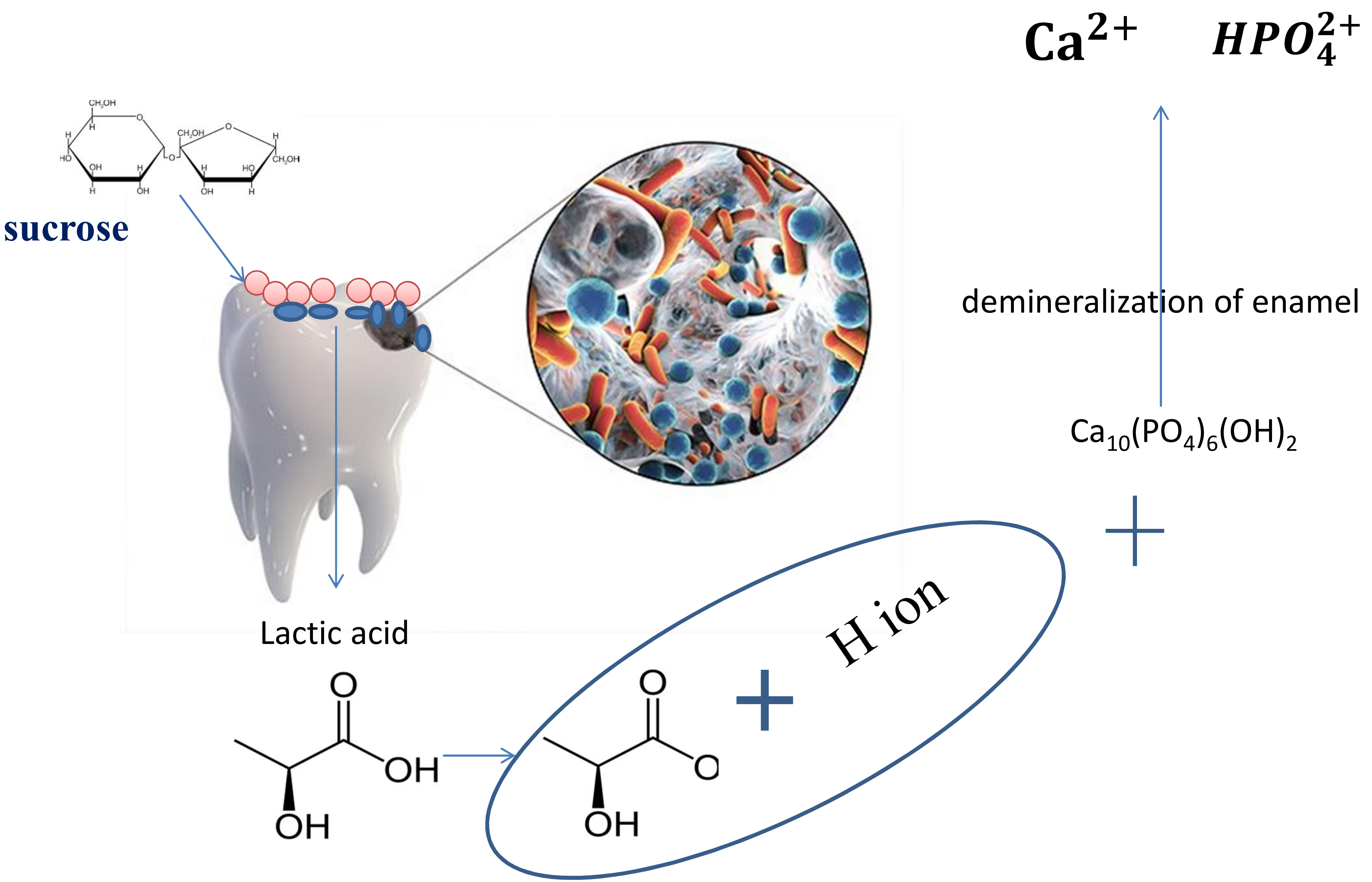
- Result from failure of the respiratory system to balance pH
- PCO_2 is the most important indicator of respiratory inadequacy
- PCO_2 levels • Normal PCO_2 fluctuates between 35 and 45 mm Hg
- Values above 45 mm Hg signal respiratory acidosis
- Values below 35 mm Hg indicate respiratory alkalosis

Respiratory Acidosis and Alkalosis

- Respiratory acidosis is the most common cause of acid-base imbalance
- Occurs when a person breathes shallowly, or gas exchange is slowed down by diseases such as pneumonia, cystic fibrosis, or emphysema
- Respiratory alkalosis is a common result of hyperventilation

pH and dental carries include

1-The relation between sucrose intake and dental carries (mechanism)



Precipitation, Acid-Base Reactions, and Tooth Decay

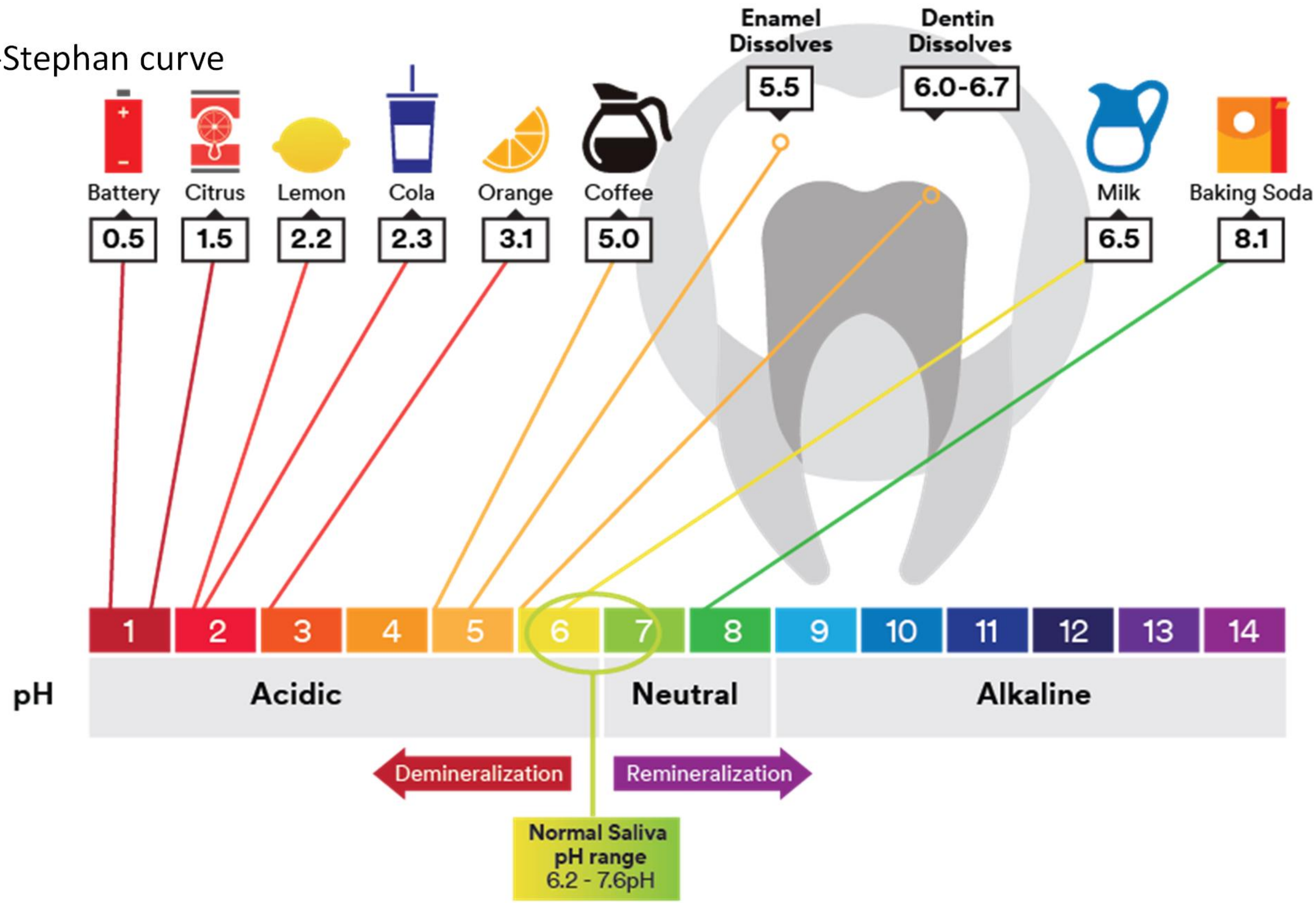
Teeth have a protective coating of hard enamel that is about 2 mm thick and consists of about 98% hydroxyapatite $Ca_5(PO_4)_3OH(s)$ Like any ionic solid surrounded by a water solution, the hydroxyapatite is constantly dissolving and reprecipitating.

$$Ca_5(PO_4)_3OH(s) \rightleftharpoons 5Ca^{2+}(aq) + 3PO_4^{3-}(aq) + OH^-(aq)$$

Saliva

enough hydroxide to keep the rate of solution and the rate of precipitation about equal

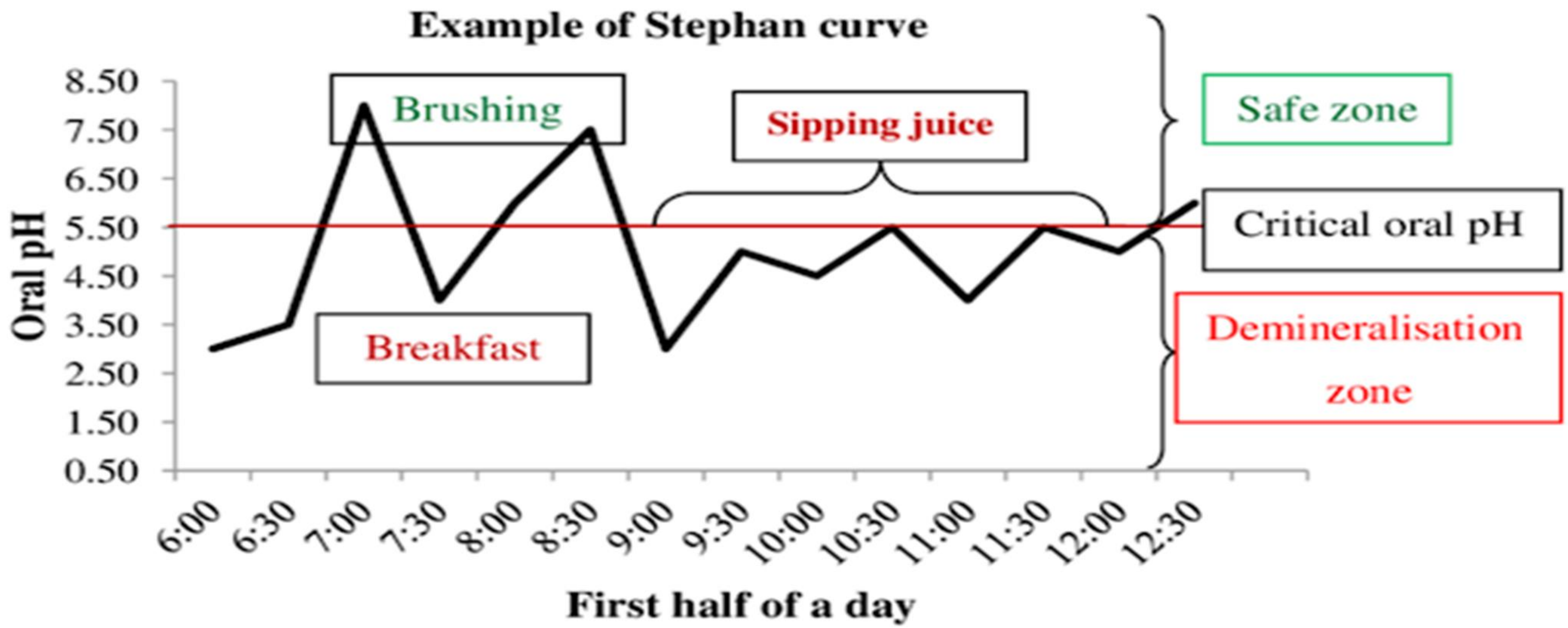
2-Stephan curve



Demineralization occurs when foods/beverages are acidic. Dentin demineralization occurs at a pH level of 6.0-6.7 and enamel demineralization occurs at a pH level of 5.5. Anything with a pH below these levels is potentially damaging to the teeth.

STEPHEN CURVE

- First described by Robert Stephen in 1943
- Stephen curve is a graph plotted on pH level against time
- This graph will show the food intake will reduce pH level in the mouth to a level bad for teeth and then rises again with time
- Stephen curve shows how the pH level at neutral in the mouth is 7 and every time it drops below 5.5 and is seemed to be critical and acid attack happens



The Mechanism of Erode and Decay

The bacteria in your mouth break down your food, especially food high in sugar, to form acids such as acetic acid and lactic acid.

These acids neutralize the hydroxide in your saliva, slowing the precipitation of enamel. The $Ca_5(PO_4)_3OH$ continues to go into solution, so there is a net loss of the protective coating on the teeth

Protecting Your Teeth

Fluoride in our drinking water and our toothpaste can help minimize the damage described above. The fluoride ion takes the place of the hydroxide ion to precipitate fluorapatite, , $Ca_5(PO_4)_3F$, a compound very similar to the

$$5Ca^{2+}(aq) + 3PO_4^{3-}(aq) + F^-(aq) \rightleftharpoons Ca_5(PO_4)_3F(s)$$