pH

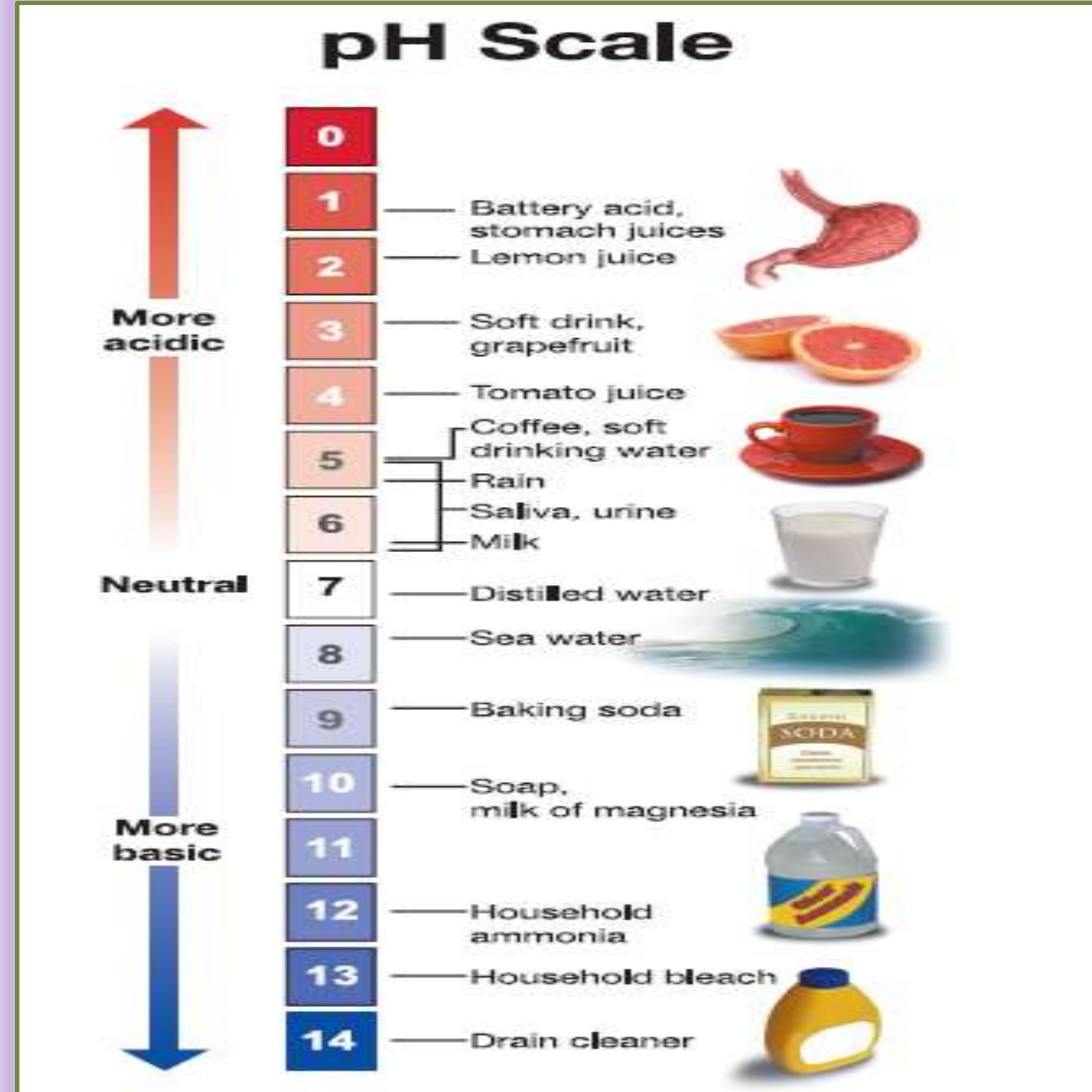
- The term “pH” refers to the measurement of the strength of an acid or base.
- “**<u>pH</u>**” comes to us from “<u>portenz of hydrogen</u>” – portenz means strength and the hydrogen’s chemical symbol is H. (p + H = pH)
- pH is measured using the pH scale which runs numerically from 0 to 14.
  - 0 – 7 is the region of acids (acidic solutions)
  - 7 is perfectly neutral (as pure water would be)
  - 7-14 is the region of bases (alkaline solutions)

The pH Scale

- The pH scale ranges from 0-14. A pH of 7 is perfectly neutral like pure water.
- As you go away from seven and count down toward zero – you are getting more and more acidic.
- As you count up from seven toward fourteen – you are getting more and more basic.
- The pH scale is based on a logarithm to the base 10 which means that each step along the pH scale represents a change in the strength of the acid or base by 10x.
- So, for acids, an acid with a pH of 3 is ten times stronger than an acid of pH 4. (10^1)
- An acid with a pH of 2 is one thousand times stronger than an acid of pH 5. ($10^3 = 10 \times 10 \times 10$)
- Bases operate by the same rule of measure.

Concentration of hydrogen ions compared to distilled water		Examples of solutions at this pH
10,000,000	pH = 0	battery acid, strong hydrofluoric acid
1,000,000	pH = 1	hydrochloric acid secreted by stomach lining
100,000	pH = 2	lemon juice, gastric acid, vinegar
10,000	pH = 3	grapefruit, orange juice, soda
1,000	pH = 4	tomato juice, acid rain
100	pH = 5	soft drinking water, black coffee
10	pH = 6	urine, saliva
1	pH = 7	"pure" water
1/10	pH = 8	sea water
1/100	pH = 9	baking soda
1/1,000	pH = 10	Great Salt Lake, milk of magnesia
1/10,000	pH = 11	ammonia solution
1/100,000	pH = 12	soapy water
1/1,000,000	pH = 13	bleaches, oven cleaner
1/10,000,000	pH = 14	liquid drain cleaner

The scale is courtesy of The Pacific Institute for the Mathematical Sciences



pH Scale
More acidic
Neutral
More basic
0
1
2
3
4
5
6
7
8
9
10
11
12
13
14
Battery acid, stomach juices
Lemon juice
Soft drink, grapefruit
Tomato juice
Coffee, soft drinking water
Rain
Saliva, urine
Milk
Distilled water
Sea water
Baking soda
Soap, milk of magnesia
Household ammonia
Household bleach
Drain cleaner
SODA

Calculating pH

- The formula for calculating ph is:

$$\mathbf{pH = -\log_{10}[H_3O^+]}$$

- The square brackets around the hydronium means that you are using the concentration of hydronium ions in the solution being measured.

Calculating pH

- pH measures the amount of hydrogen ions released by the acid when placed in a solution. (Recall the link between dissociation and acid/base strength from previous lecture.)
- When an acid is placed in a solution, it releases the hydrogen (H^+) ions and these ions attach themselves to the water molecules (H_2O) to make the **hydronium ion** (**H_3O^+**). The more hydronium ions that are made – the more H^+ ions that must have been released – the stronger the acid.

$$\mathbf{H^+ + H_2O \rightarrow H_3O^+}$$

- This means that to measure the pH of an acid, we are actually measuring the number of hydronium ions rather than the number of hydrogen ions because the hydrogen ions quickly form hydroniums.

pH Measuring Electrode

- Purpose is to Develop a Millivolt Potential Directly Proportional to the Free Hydrogen Ion Concentration in an Aqueous Solution
  - Process Effects
    - **High Temperature**
      - Faster Response / Lower Impedance
      - Accelerates Aging, Lithium Ions Leached from Membrane
        - » Short Span
    - **Low Temperature**
      - Slower Response / Higher Impedance
    - **Measurement > 10.0 pH**
      - Alkaline / Sodium Ion Error
    - **Coatings**
      - Slower Response
      - Increase Zero Offset
    - **< 50% Water**
      - Dehydration
    - **Steam Sterilization**
      - Dehydration
      - Ag/AgCl Dissolves from Silver Reference Element

Using pH meters

1. Always rinse pH meter in distilled water prior to placing it in a solution (buffer or otherwise)
2. Place the pH meter in a buffer with about the same pH as that of your solution (4, 7, or 10)
3. Turn on the pH meter only when in solution
   Start with with buffer 7. Hit “cal” once.
   Wait upto a minute until it automatically sets
   Rinse, dry, and place in second buffer (4/10)
   Hit “cal” once. Wait until it automatically sets
   There is no need to use “read”.
4. Measure the pH of your solution

pOH

- The pH measures the strength of an acid but it does not do the same for a base.
- A base is a hydrogen ion acceptor so when you put it in water – it doesn't form the hydronium ion – it forms the hydroxide ion (OH^-) because it takes an H^+ from the water.

$$\mathbf{H_2O - H^+ \rightarrow OH^-}$$

- This means that the accurate way to measure the strength of a base is to measure the number of hydroxide ions because the base molecules steal H^+ ions from the water and quickly turn the water molecules into hydroxide ions (OH^-).

Calculating the pOH

- The formula for pOH is:

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

- Again, the squared brackets tells you the value within them is the concentration of the hydroxide ions in the solution being measured.
- And again, the calculator should be the toughest part of figuring out how to solve these problems.

The Henderson-Hasselbalch Equation

- A buffer must be prepared from a conjugate acid-base pair in which the **K_a of the acid is approximately equal to the desired H_3O^+ concentration**.
  - To illustrate, consider a buffer of a weak acid HA and its conjugate base A^-.

According to the Brønsted-Lowry theory of acids and bases, an acid (HA) is capable of donating a proton (H^+) and a base (B) is capable of accepting a proton. After the acid (HA) has lost its proton, it is said to exist as the conjugate base (A^-). Similarly, a protonated base is said to exist as the conjugate acid (BH^+).

The dissociation of an acid can be described by an equilibrium expression:

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

Consider the case of acetic acid (CH_3COOH) and acetate anion (CH_3COO^2):

$$\underset{\text{Acetic acid}}{CH_3COOH} \qquad \underset{\text{Acetate}}{CH_3COO^- + H^+}$$

Acetate is the conjugate base of acetic acid. Acetic acid and acetate are a conjugate acid/base pair. We can describe this relationship with an equilibrium constant:

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

In this simulation, we will use K_A for the acid dissociation constant. Taking the negative log of both sides of the equation gives:

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

This can be rearranged:

$$-\log K = -\log[H^+] + (-\log \frac{[A^-]}{[HA]})$$

By definition, pK_A = $-\log K_A$ and pH = $-\log[H^+]$, so

$$pK = pH - \log \frac{[A^-]}{[HA]}$$

This equation can then be rearranged to give the Henderson-Hasselbalch equation:

$$pH = pK + \log \frac{[A^-]}{[HA]} = pK + \log \frac{[\text{conjugate base}]}{[\text{conjugate acid}]}$$

The Henderson-Hasselbalch equation can be used to prepare buffer solutions and to estimate charges on ionizable species in solution, such as amino acid side chains in proteins. Caution must be exercized in using this equation because pH is sensitive to changes in temperature and salt concentration in the solution being prepared.

Buffers

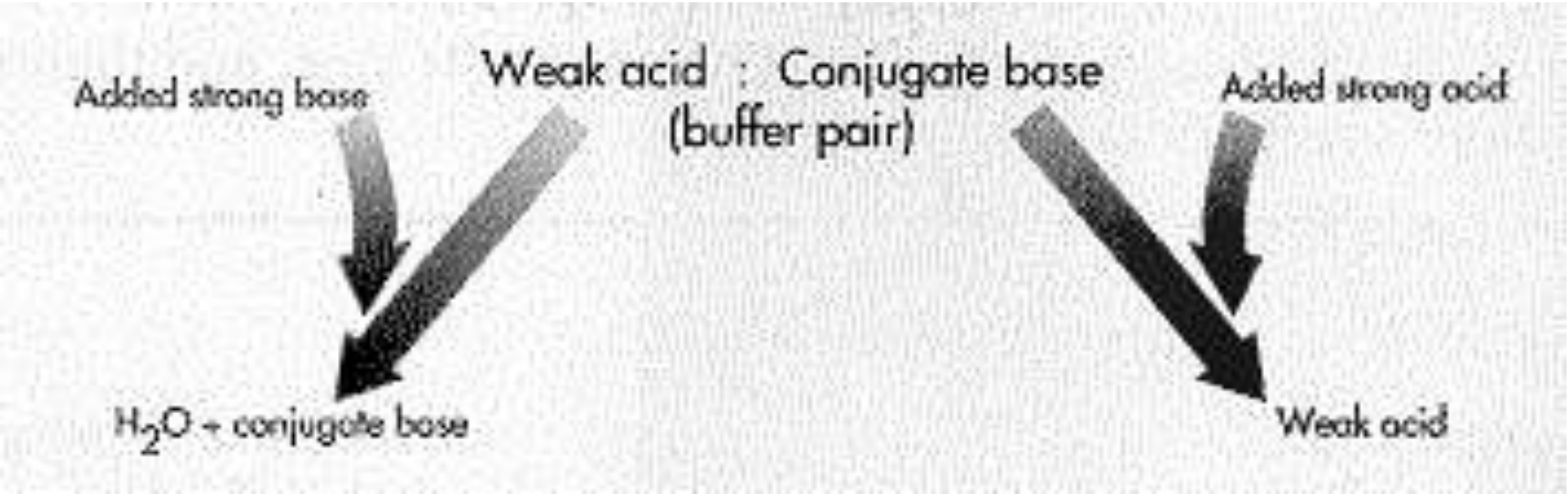
- Definition
  - Chemical Substances
  - Present in Body Fluids: Where these Reactions
- Occur
  - Prevent a Sharp/ Drastic Change
    - pH of that Fluid
    - When an Acid or Base is Added / Introduced to it
  - Normal Body Metabolism
    - Tends to Produce More H+ (acids) than OH- (bases)
    - Making the Blood More Acidic than Basic.

Buffers

- Buffers Pairs
- 2 Substances: Acid & Base
- Job: Buffer the Blood
- Main Buffer Pair
  - Sodium Bicarbonate ($NaHCO_3$ or Ordinary Baking Soda)

  - Carbonic Acid (H_2CO_3).
- Without Buffers
- – Blood Would Become too Acidic
- – We Would Die.

Mechanisms by which Buffer Operate



Example:

CH3COOH + CH3COO- + OH- = 2CH3COO- + H2O (you get more conjugate base)

CH3COOH + CH3COO- + H+ = 2CH3COOH (you get more weak acid)

Results of Not Buffering

- Place a Strong Acid or Strong Base in Water
  - pH will Change Dramatically.
  - Add HCl to Water: $HCl + H2O \rightarrow H3O+ + Cl-$.
- Proton (H+) from the HCl
  - Bind With Water Molecules (Neutral)
  - Form H3: Raise the concentration of H+ In Solution.
- Solution Now More Acidic: Big Drop in the pH reading
- Effect is Immediate: Not Long-Lasting

Buffering Action of Sodium Bicarbonate

- $NaHCO_3$ Buffers HCl
- Strong Acid Becomes Weaker One: Carbonic Acid (H + HCO_3)
- HCl Releases More H+ Ions in Water (Dissociates)
- More H+ Ions in Solution: More Acidic
- Adding $NaHCO_3$ (Buffer): Resulting Carbonic Acid Releases
- Fewer H+ Ions in Water
- Solution Less Acidic

Buffering Action of Carbonic Acid

- Carbonic Acid Buffers NaOH
- Strong Base Becomes a Weaker One: Sodium
- Bicarbonate (Na + HCO3)
- NaOH Releases More Hydroxide Ions in Water
- (Disassociates)
- More OH Ions in Solution: More Alkaline
- Adding H2CO3 (Buffer): Resulting HCO3 Releases
- Less OH- in Water
- Solution Less Alkaline

pH of Body fluids

- pH
  - Number that Indicates the Hydrogen Ion (H+) Concentration of a Fluid
- Presented: Logarithmic Scale of 0 – 14.
- Each Unit (or #): 10 Fold Increase or Decrease.
- Lower the #
  - Higher the H+ Concentration
  - More Acidic Solution
- Higher the #
  - Lower the H+ Concentration
  - More Alkaline Solution

pH Scale

– pH 7

- Indicates Neutrality (pure water)
- Contains: Equal Amount of Hydrogen (H^+) & Hydroxide ions(OH^-)

– pH Higher than 7

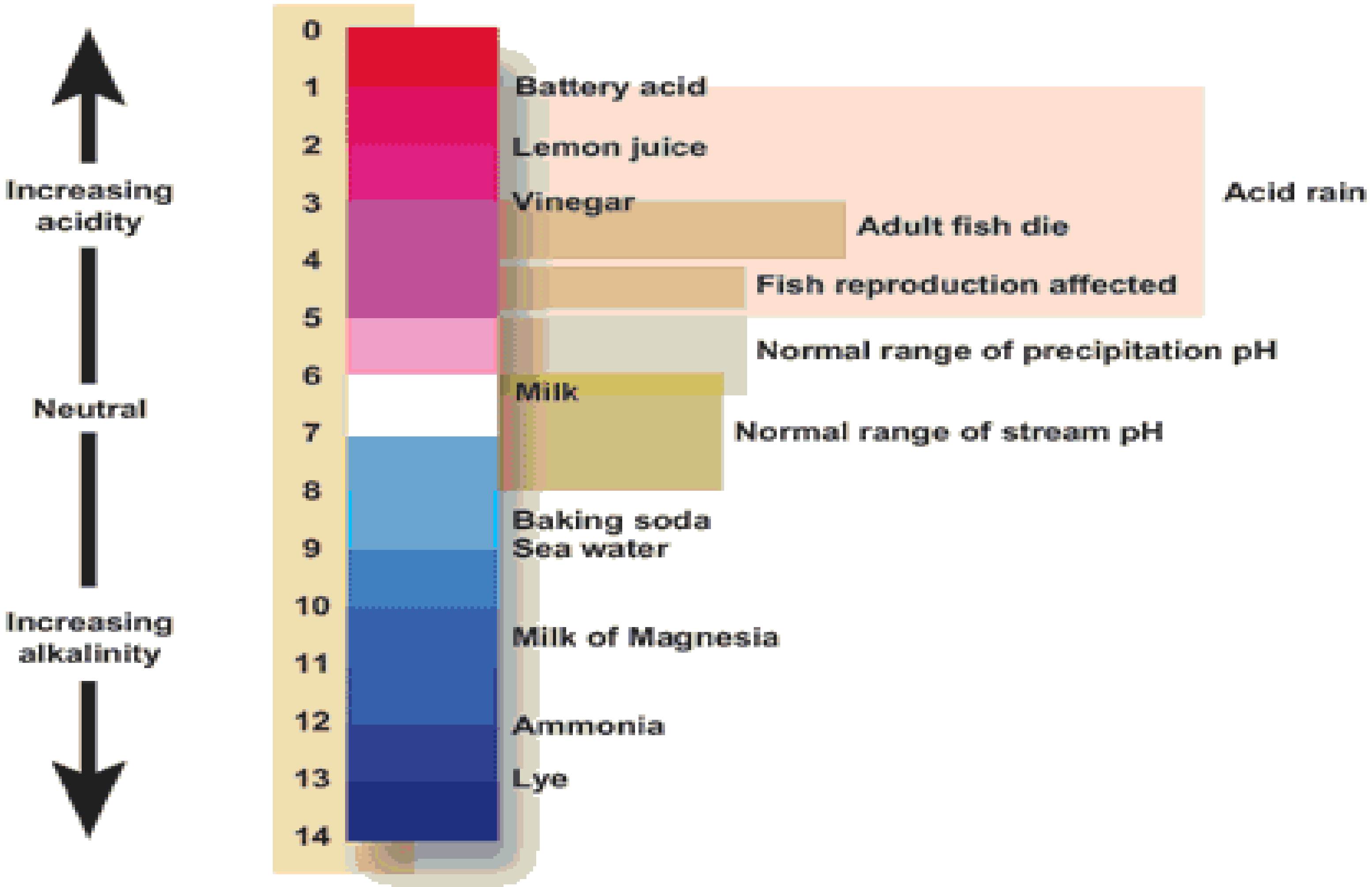
- Indicates Alkalinity or Base
- Contains: Less H^+ Ions than OH^- Ions

– pH Less than 7

- Indicates Acidity
- Contains: Higher Concentration / More H^+ Ions than OH^- Ions

pH Range of Blood

- Normal Arterial Blood pH—about 7.45
- Normal Venous Blood pH—about 7.35
- Vary Narrow Range!!



Increasing acidity
Neutral
Increasing alkalinity
0
1
2
3
4
5
6
7
8
9
10
11
12
13
14
Battery acid
Lemon juice
Vinegar
Adult fish die
Fish reproduction affected
Acid rain
Normal range of precipitation pH
Milk
Normal range of stream pH
Baking soda
Sea water
Milk of Magnesia
Ammonia
Lye