(Renal Physiology 9)

Acid-Base Balance 1

Basics of Acid Base

Ahmad Ahmeda

aahmeda@ksu.edu.sa

Cell phone: 0536313454

Learning Objectives:

- Define: acid and base.
- Explain what is meant by strong and weak acids and bases.
- List and identify the names/formulas for the common strong acids and strong bases.
- To explain the role of Henderson-Hasselbalch equation in acid-base regulation.

Acid – Base Balance

- Acid – Base balance (*a.k.a.* **pH HOMEOSTASIS**) one of the essential functions of the body.

- When discussing acid - base balance, we are normally concerned with regulation of **H^+ ion** balance (although HCO_3^- plays a vital role in this balance).

- To avoid disturbances in [H^+], and to maintain its homeostasis, the amount generated / taken in **MUST EQUAL** the amount secreted.

Acid – Base Balance

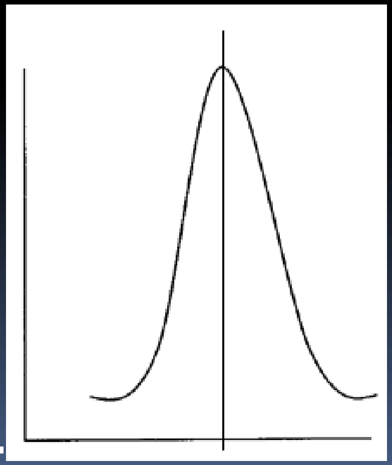
Why is control of [H^+] so important?

- Highly reactive chemical species (protons).
  - combine easily with negatively charged ions and bases.

- Precise [H^+] regulation is vital because activity of almost all enzyme systems / proteins (inc. ion channels) influenced by pH **(*e.g.* hydrogen bonding and charge on proteins altered by pH ∴ tertiary structure and function affected.)**

- Most enzymes function optimally at pH ~ 7.4 (except gastric enzymes).

- Acid-base imbalances can cause cardiac arrhythmias and abnormal neuronal excitation.



Acid – Base Balance

Definitions (Bronsted-Lowry)

- **ACIDS** – Molecules containing hydrogen atoms that can release (**DONATE**) H^+ into solution (*e.g.* $HCl \Leftrightarrow H^+ + Cl^-$).

- **STRONG** acids – all their H^+ is dissociated completely in H_2O.
- **WEAK** acids – dissociate partially in H_2O and are efficient at preventing pH changes.

- **BASES** (*a.k.a.* alkalis) – ions or molecules that can **ACCEPT** H^+ (*e.g.*, $HCO_3^- + H^+ \Leftrightarrow H_2CO_3$).

- **STRONG** bases – dissociate easily in H_2O and quickly bind H^+.
- **WEAK** bases – accept H^+ more slowly (*e.g.*, HCO_3^- and NH_3)

  Proteins in body function as weak bases as some constituent **AMINO ACIDS** have net negative charge and attract H^+ (*e.g.* **HAEMOGLOBIN**).

Free hydrogen ions are extremely unstable. Therefore, for any acid and any base, the equilibrium established is:

$$AH + B \rightleftharpoons A^- + BH^+$$

Where AH is an acid and A is its conjugate base and B is a base and BH is its conjugate acid.

In other words, every acid has a conjugate base associated with it, and vice versa.

Water usually is amphoteric compound because it can act like an acid or base depend on the added compound (acidic or basic)

When water behaves as a base, it accepts H^+ and forms a hydronium ion; H_3O^+. When it behaves as an acid, it loses a proton, and forms a hydroxide ion; OH

pH and pK_a:

The **pH** of a solution is a measure of the acidity of the solution. It is defined as:

$$\mathrm{pH} = -\log_{10}\left([\mathrm{H_3O^+}]\right)$$

Where [H_3O+] is the concentration of hydronium ions in the solution.

The pH of a solution depends on two things:

1. The concentration of the solution, if we have two solutions of the same acid, the more concentrated solution will have more free H_3O^+ ions and therefore a lower pH.

2. The acid in question, if we have two equally concentrated solutions of acids, the solution of a strong acid will have a lower pH than that of a weak acid, because it is more fully dissociated and therefore produces more H_3O^+ ions. HCl, for example, is completely dissociated.

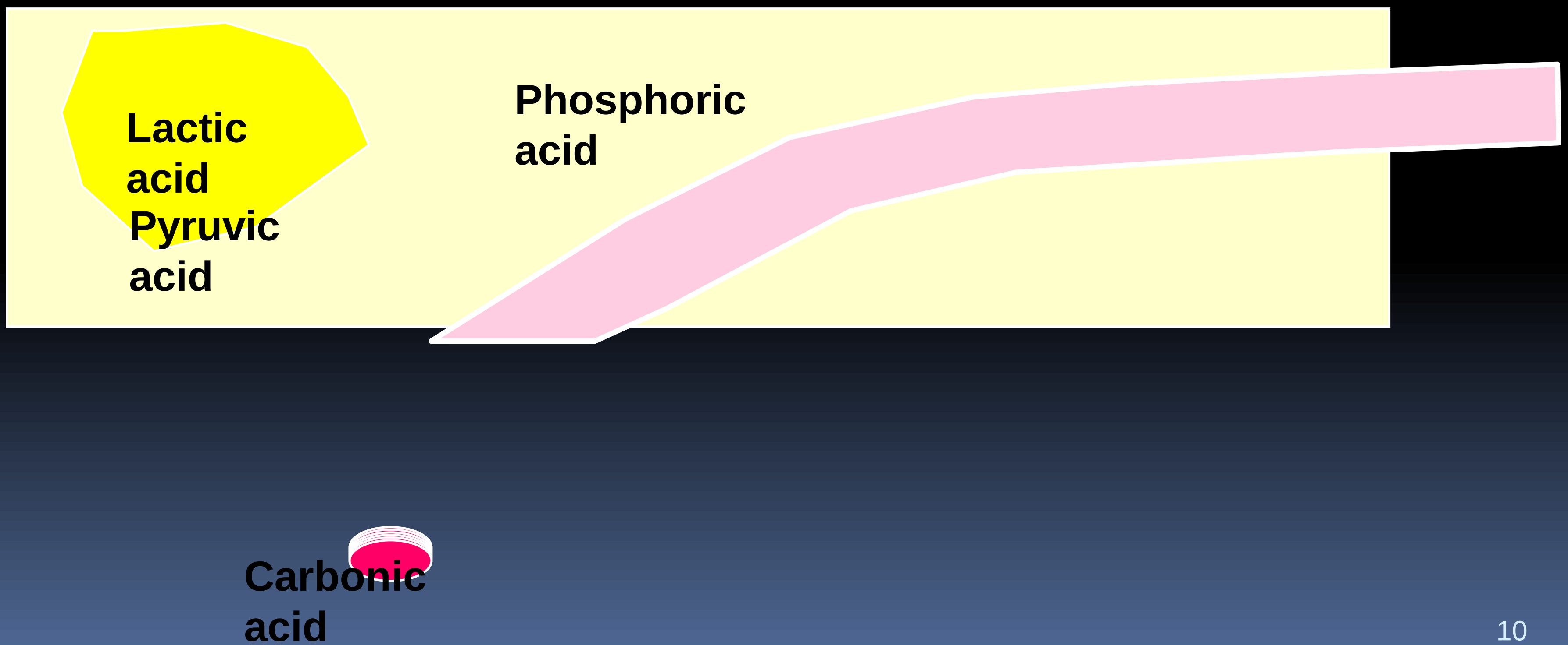
Therefore, we see that pH does not measure the strength of an acid, but the acidity of a given solution.

The pH of water is 7. This means that a solution of pure water has a 10^{-7} mol dm^{-3} of hydronium ions.

Physiologically important acids include:

- **Carbonic acid (H_2CO_3)**
- **Phosphoric acid (H_3PO_4)**
- **Pyruvic acid ($C_3H_4O_3$)**
- **Lactic acid ($C_3H_6O_3$)**

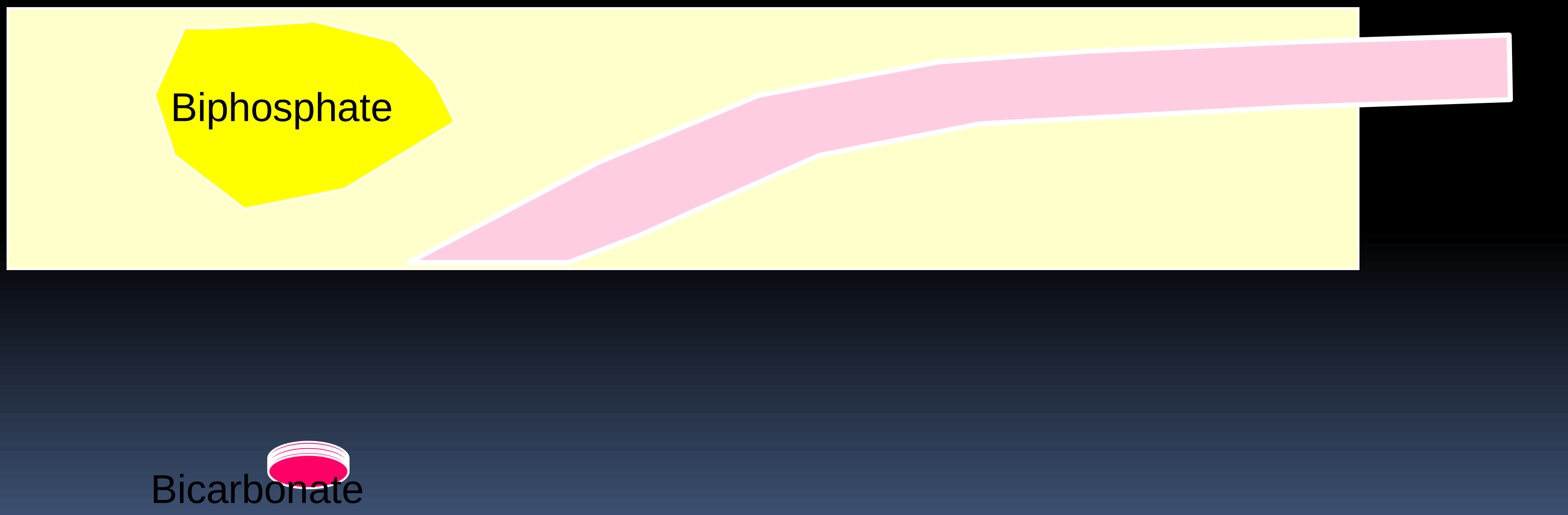
These acids are dissolved in body fluids



Physiologically important bases include:

Bicarbonate (HCO_3^-)

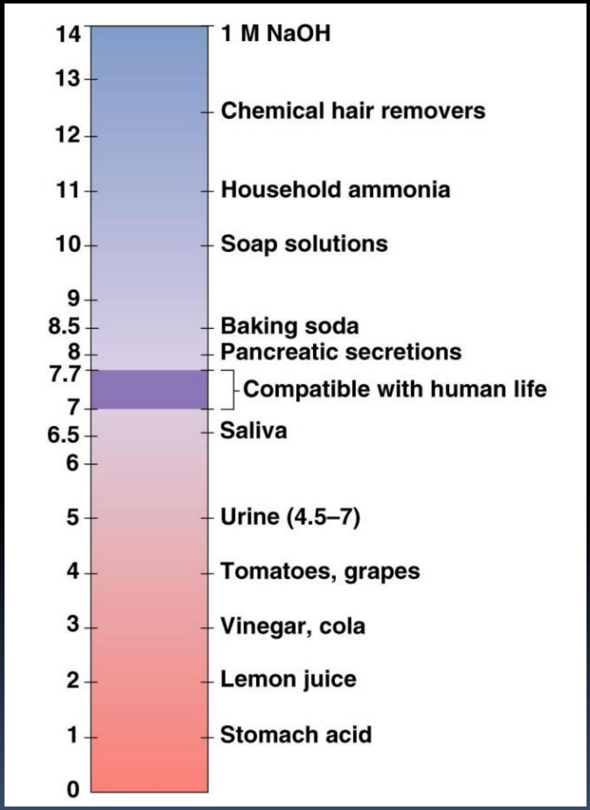
Biphosphate (HPO_4^{-2})



Acid – Base Balance

pH Scale (Sørensen, 1909)

- Relative to other ions, [H^+] of body fluids kept **VERY LOW**
  e.g., ECF [Na^+] ≈ 145 mM/L, ECF [H^+] ≈ 0.00004 mM/L (40nM)
  (~ 3.5 million fold difference).



- Because [H^+] so low, easier to express [H^+] on a logarithmic scale ⇒ pH units.

$$\mathbf{pH = \log \frac{1}{[H^+]} = -\log [H^+]}$$

- Normal pH = -log [0.00000004] M
  = 7.4

Acid – Base Balance

pH Scale (Sørensen, 1909)

table 2.5 Hydrogen Ion Concentrations and pH

Grams of H^+ per Liter	pH	
0.00000000000001	14	↑
0.0000000000001	13	
0.000000000001	12	Increasingly basic
0.00000000001	11	
0.0000000001	10	
0.000000001	9	
0.00000001	8	Neutral–neither
0.0000001	7	acidic nor basic
0.000001	6	
0.00001	5	
0.0001	4	
0.001	3	Increasingly acidic
0.01	2	
0.1	1	
1.0	0	↓

Note that a change of 1 pH unit = 10x change in $[H^+]$ ($\log_{10}$ scale)

Acid – Base Balance

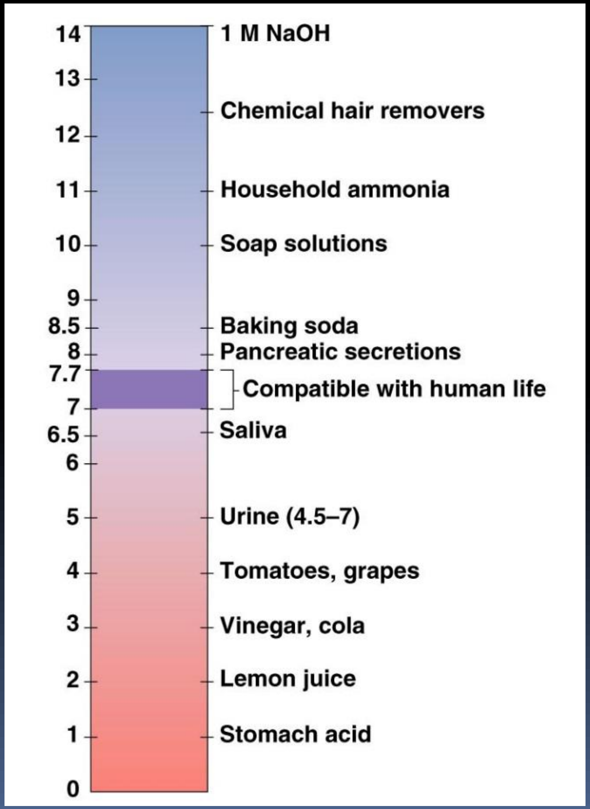
pH Scale (Sørensen, 1909)

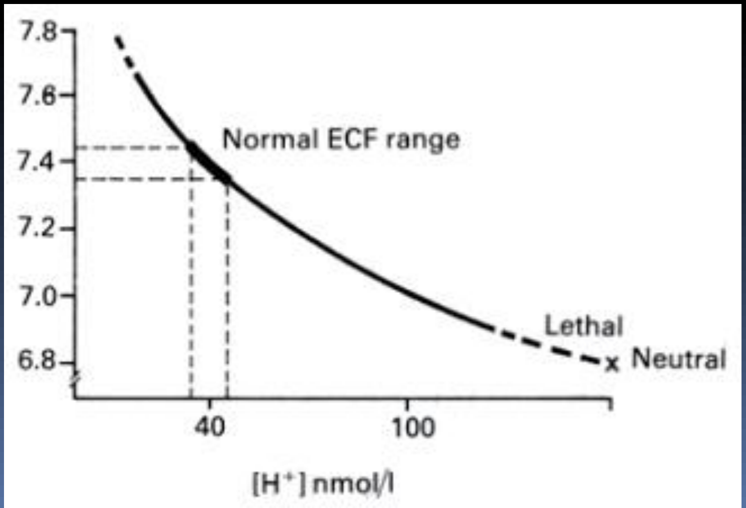
- pH **INVERSELY** related to [H^+], *i.e.*

  as [H^+] ↑, pH falls – acidosis (below 7.35)

  as [H^+] ↓, pH increases – alkalosis (above 7.45)

- Normal **BLOOD** pH range for adults = **7.35 – 7.45** maintained by chemical buffer systems, kidneys and lungs.
  - **DEATH** likely if pH ↑7.8 or ↓6.8.





Acid – Base Balance

pH Scale (Sørensen, 1909)

- However, there are a range of pH values within different body fluids - dependent on function.

Approximate pH Values of Various Body Fluids

Compartment	pH
Gastric secretions (under conditions of maximal acidity)	0.7
Lysosome	5.5
Chromaffin granule	5.5
Neutral H_2O at 37°C	6.81
Cytosol of a typical cell	7.2
Cerebrospinal fluid	7.3
Arterial blood plasma	7.4
Mitochondrial inner matrix	7.5
Secreted pancreatic fluid	8.1

pH and $[H^+]$ of Body Fluids

	H^+ Concentration (mEq/L)	pH
Extracellular fluid		
Arterial blood	4.0×10^{-5}	7.40
Venous blood	4.5×10^{-5}	7.35
Interstitial fluid	4.5×10^{-5}	7.35
Intracellular fluid	1×10^{-3} to 4×10^{-5}	6.0 to 7.4
Urine	3×10^{-2} to 1×10^{-5}	4.5 to 8.0
Gastric HCl	160	0.8

Sources of H^+

- The body generally **PRODUCES** more acids than bases.

1) Cellular aerobic metabolism produces **15,000 mmol** CO_2/day.

$$CO_2 + H_20 \leftrightarrow H_2CO_3 \leftrightarrow H^+ + HCO_3^-$$

(volatile acid)

  - Normally all volatile acid **excreted by the lungs**.

2) **DIET** – incomplete metabolism of carbohydrates (lactate) lipids (ketones) and proteins (H_2SO_4, H_3PO_4) generates **fixed (non-volatile) acids** – ~50 -100 mEq per day.

- In order to maintain balance, acids need to be **BUFFERED** and/or **EXCRETED**.

Notes:

Moles x valence = equivalents.

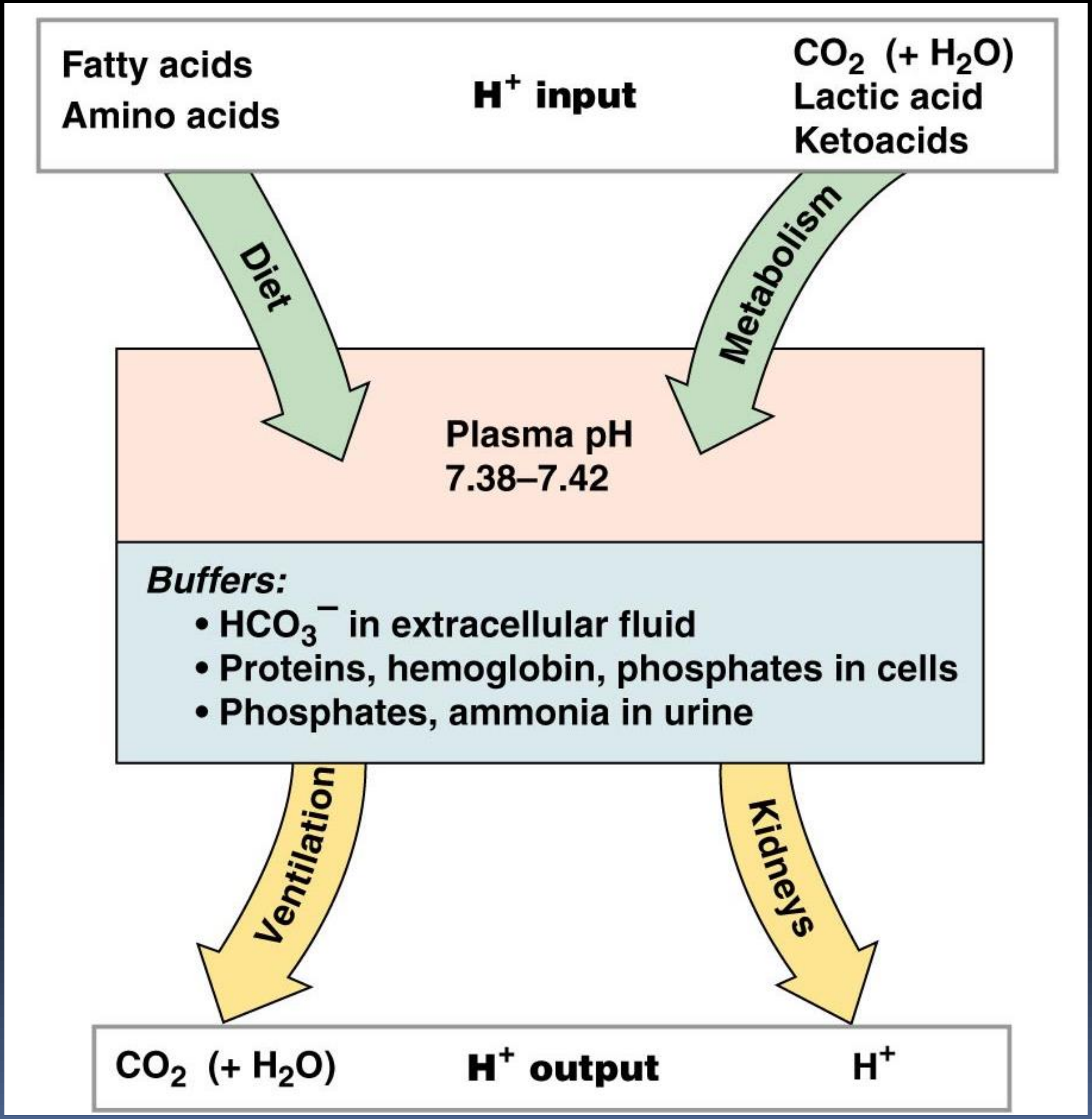
By definition one equivalent (or equivalent weight) of a substance is the amount of that substance which supplies or consumes one mol of reactive species. In acid-base chemistry the reactive species is the hydrogen ion (H^{1+}) while in oxidation-reduction chemistry the reactive species is the electron. For example, in the following two reactions the equivalent weight of H_2SO_4 would be 49 grams or 0.5 mol in the first reaction but 98 grams or 1 mol in the second. On the other hand, sodium hydroxide has the same equivalent weight in both reactions, one mol or 40 grams.

(1) $H_2SO_4 + 2NaOH$ ------> $Na_2SO_4 + 2H_2O$

(2) $H_2SO_4 + NaOH$ ------> $NaHSO_4 + H_2O$

In the first reaction one mol of H_2SO_4 supplies 2 mols of H^{1+} to NaOH, therefore, one-half mol of H_2SO_4 or 49 grams is one equivalent. The conditions are different in the second reaction because sulfuric acid only "loses" one hydrogen so the equivalent weight of sulphuric acid is one mol or 98 grams. However, sodium hydroxide behaves the same in both reactions, that is, one mol of sodium hydroxide always "consumes" one mol of H^{1+}, so its equivalent weight remains the same at one mol or 40 grams.

Hydrogen and pH Balance in the Body



Catabolism of sulphur containing AA gives H_2SO^4

Catabolism of phospholipids/ phosphoproteins give H_3PO^4

How is $[H^+]$ Controlled?

- Three systems involved;

  1) **BUFFERS** – first defence

     - second – to - second regulation of $[H^+]$

  2) Excretion of CO_2 ($\downarrow H_2CO_3$) by **LUNGS** (removal of volatile acid) – second defence

     - regulation in minutes - to - hours

  3) Excretion of H^+ ($\uparrow HCO_3^-$) by **KIDNEYS** (fixed acids)

     - third defence
     - regulation over several hours to days
     - slowest, but most **POWERFUL**, of body's acid-base regulatory systems.

- Relative concentrations of CO_2 and HCO_3^- in plasma / ECF determine pH (**HENDERSON-HASSELBALCH** equation) .

(show the relationship between pH, hydrogen ion conc. and the ratio of buffer membrane in a solution)

HENDERSON-HASSELBALCH equation

$$pH = pK' + \frac{\log [HCO_3^-]}{sPCO_2}$$

pH is the negative logarithm of H^+ in mol/L.
pK' is negative logarithm of overall dissociation constant for the reaction = 6.1 in health.
s is solubility of CO_2 in solution = 0.03 at 37ºC

$$pH = 6.1 + \frac{\log [HCO_3^-]}{0.03 \times PCO_2}$$

In health, $[HCO_3^-]$ = 24 mmol/L & PCO_2 = 40 mm Hg

7.4 = 6.1 + 1.3

Note: pKa tells us how acidic (or not) a given hydrogen atom in a molecule is. pH tells us how acidic a solution is.

What happen to the pH using H-H

In case if the HCO_3 in Plasma remains normal

- If Pco_2 increased, the ratio of $[HCO_3]P/ 0.03 Pco_2$ will decrease which lead to acidosis.
- If Pco_2 decrease, the ratio will increase and pH will increase causing alkalosis.

In case the Pco_2 remains normal

- Increase bicarbonate in plasma causes an increase in the ratio which leads to alkalosis.
- Decrease in bicarbonate in plasma causes a decrease in the ration which leads to acidosis.

Thanks