CHAPTER 7

Carbon Dioxide Transport and Acid-Base Balance
Carbon Dioxide Transport

• In plasma:
  – Carbamino compound (bound to protein)
  – Bicarbonate
  – Dissolved CO$_2$
CO₂ is converted to HCO₃⁻ at the tissue sites. Most of the CO₂ that is produced at the tissue cells is carried to the lungs in the form of HCO₃⁻.

Fig. 7-1. How CO₂ is converted to HCO₃⁻ at the tissue sites. Most of the CO₂ that is produced at the tissue cells is carried to the lungs in the form of HCO₃⁻.

\[
P_{CO_2} \text{ Directly Affects } \frac{H_2CO_3}{HCO_3^-}\text{ Levels in Plasma}
\]

\[
H_2CO_3 = P_{CO_2} \times 0.0301
\]
Carbon Dioxide Transport

• In red blood cells:
  – Dissolved CO$_2$
  – Carbamino-Hb
  – Bicarbonate
CARBON DIOXIDE ELIMINATION AT THE LUNGS
How HCO₃⁻ Is Transformed into CO₂

Fig. 7-2. How HCO₃⁻ is transformed back into CO₂ and eliminated in the alveoli.

\[ \text{H}_2\text{CO}_3 = P_{\text{CO}_2} \times 0.0301 \]
# Carbon Dioxide (CO₂) Transport Mechanisms

<table>
<thead>
<tr>
<th>CO₂ Transport Mechanisms</th>
<th>Approx. % of Total CO₂ Transported to the Lungs</th>
<th>Approx. Quantity of Total CO₂ to the Lungs (ML/MIN)</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>IN PLASMA</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Carbamino compound</td>
<td>1</td>
<td>2</td>
</tr>
<tr>
<td>Bicarbonate</td>
<td>5</td>
<td>10</td>
</tr>
<tr>
<td>Dissolved CO₂</td>
<td>5</td>
<td>10</td>
</tr>
<tr>
<td><strong>IN RED BLOOD CELLS</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Dissolved CO₂</td>
<td>5</td>
<td>10</td>
</tr>
<tr>
<td>Carbamino-Hb</td>
<td>21</td>
<td>42</td>
</tr>
<tr>
<td>Bicarbonate</td>
<td>63</td>
<td>126</td>
</tr>
<tr>
<td>Total</td>
<td>100</td>
<td>200</td>
</tr>
</tbody>
</table>

Table 7-1
Carbon Dioxide Dissociation Curve

• Similar to the oxygen dissociation curve, the loading and unloading of CO\textsubscript{2} in the blood can be illustrated in graphic form.
Carbon Dioxide Dissociation Curve

Fig. 7-3. Carbon dioxide dissociation curve.
Carbon Dioxide Dissociation Curve

Fig. 7-4. Carbon dioxide dissociation curve. An increase in the PCO$_2$ from 40 mm Hg to 46 mm Hg raise the CO$_2$ content by about 5 vol.%. PCO$_2$ changes have a greater effect on CO$_2$ content levels than PO$_2$ changes on O$_2$ levels.
Carbon Dioxide Dissociation Curve

Fig. 7-5. Carbon dioxide dissociation curve at two different oxygen/hemoglobin saturation levels (SaO₂ of 97% and 75%). When the saturation of O₂ increases in the blood, the CO₂ content decreases at any given PCO₂. This is known as the Haldane effect.
Fig. 7-6. Comparison of the oxygen and carbon dioxide dissociation curves in terms of partial pressure, content, and shape.
ACID-BASE BALANCE AND REGULATION
• Nearly all biochemical reactions in the body are influenced by the acid-base balance of their fluid environment

• Normal arterial pH range is 7.34 to 7.45
Acid-Base Balance and Regulation

- pH > 7.45 = alkalosis
- pH < 7.35 = acidosis
• Most $H^+$ ions in the body originate from:
  1. Breakdown of phosphorous-containing proteins (phosphoric acid)
  2. Anaerobic metabolism of glucose (lactic acid)
  3. Metabolism of body fats (fatty and ketone acids)
  4. Transport of $CO_2$ in the blood as $HCO_3^-$ liberates $H^+$ ions
H$^+$ and HCO$_3^-$

- Under normal conditions, both the H$^+$ and HCO$_3^-$ ion concentrations in the blood are regulated by the following three major systems:
  - Chemical buffer system
  - Respiratory system
  - Renal system
The Chemical Buffer System

• Responds within a fraction of a second to resist pH changes
  – Called the first line of defense
The Chemical Buffer System

• System is composed of:
  1. Carbonic acid-bicarbonate buffer system
  2. Phosphate buffer system
  3. Protein buffer system
• Acts within one to three minutes by increasing or decreasing the breathing depth and rate to offset acidosis or alkalosis, respectively.
The Renal System

- Body’s most effective acid-base balance monitor and regulator
- Renal system requires a day or more to correct abnormal pH concentrations
• Note:
  – To fully appreciate acid-base balance, and how it is normally regulated, a fundamental understanding of acids and bases, and their influences on pH, is essential.
THE BASIC PRINCIPLES OF ACID-BASE REACTIONS AND pH
• Similar to salts, acids and bases are electrolytes
• Thus, both acids and bases can:
  – Ionize and dissociate in water
  – Conduct an electrical current
• Acids are sour tasting, can react (dissolve) with many metals, and can “burn” a hole through clothing

• An acid is a substance that releases hydrogen ions \([H^+]\) in measurable amounts
Acids

• Because a hydrogen ion is only a hydrogen nucleus proton, acids are defined as proton donors.
• Thus, when acids dissolve in a water solution, they release hydrogen ions (protons) and anions.
• Acidity of a solution is directly related to the concentration of protons.
• Anions have little or no effect on the acidity.
• Thus, the acidity of a solution reflects only the free hydrogen ions, not those bound to anions.
• For example, *hydrochloric acid* (HCl), the acid found in the stomach and works to aid digestion, dissociates into a proton and a chloride ion:

\[
\text{HCl} \quad \Delta \quad \text{H}^+ \quad \text{Cl}^- \\
\text{proton} \quad \text{anion}
\]
• Other acids in the body include acetic acid (HC₂H₃O₂), often abbreviated as [HAc], and carbonic acid (H₂CO₃)
• Molecular formula for common acids is easy to identify because it begins with the hydrogen ion
• Acidity of a solution reflects only the free hydrogen ions
  – Not the hydrogen ions still combined with anions.

• Thus, strong acids, which dissociate completely (i.e., they liberate all the H⁺) and irreversibly in water, dramatically change the pH of the solution.
Strong Acids

• For example, if 100 hydrochloric (HCl) acid molecules were placed in 1 mL of water, the hydrochloric acid would dissociate into 100 $\text{H}^+$ and 100 $\text{Cl}^-$ ions.
• There would be no undissociated hydrochloric acid molecules in the solution.
Weak Acids

• Do not dissociate completely in a solution and have a much smaller effect on pH

• Although weak acids have a relatively small effect on changing pH levels, they have a very important role in resisting sudden pH changes.
Examples of weak acids are carbonic acid (H$_2$CO$_3$) and acetic acid (HC$_2$H$_3$O$_2$). If 100 acetic acid molecules were placed in 1 mL of water, the following reaction would occur:

$$100 \text{ HC}_2\text{H}_3\text{O}_2 \Delta 90 \text{ HC}_2\text{H}_3\text{O}_2 + 10 \text{ H}^+ + 10 \text{ C}_2\text{H}_3\text{O}_2^-$$

(Hydrogen ions) (Acetate ions)
Because undissociated acids do not alter the pH, the acidic solution will not be as acidic as the HCl solution discussed above.

The dissociation of acetic acid can be written as follows:

\[ \text{HC}_2\text{H}_3\text{O}_2 \rightleftharpoons \text{H}^+ + \text{C}_2\text{H}_3\text{O}_2^- \]
• Using this equation, it can be seen that:
  – When $H^+$ (released by a strong acid) is added to the acetic acid solution
  – Equilibrium moves to the left as some of the additional $H^+$ bonds with $C_2H_3O_2^-$ to form $HC_2H_3O_2$
Weak Acids

• On the other hand, when a strong base is added to the solution (adding additional OH\(^-\) and causing the pH to increase), the equilibrium shifts to the right.

• This occurs because the additional OH\(^-\) consumes the H\(^+\).
Weak Acids

- Cause more $\text{HC}_2\text{H}_3\text{O}_2$ molecules to dissociate and replenish the $\text{H}^+$
- Weak acids play a very important role in the chemical buffer systems of the human body
Bases

- Bases are proton acceptors
- Bases taste bitter and feel slippery
- A base is a substance that takes up hydrogen ions \([H^+]\) in measurable amounts
• Common inorganic bases include the hydroxides
  – Magnesium hydroxide (milk or magnesia) and sodium hydroxide (lye)
• Similar to acids, when dissolved in water, hydroxides dissociate into hydroxide ions (OH\(^{-}\)) and cations
• For example, ionization of sodium hydroxide (NaOH) results in a hydroxide ion and a sodium ion
• Liberated hydroxide ion then bonds, or accepts, a proton present in the solution
• Reaction produces water and, at the same time, decreases the acidity [H\(^+\) concentration] of the solution
• The solution:

\[ \text{NaOH} \quad \Delta \quad \text{Na}^+ + \text{OH}^- \]

cation \quad \text{hydroxide ion}

• And then:

\[ \text{OH}^- + \text{H}^+ \quad \Delta \quad \text{H}_2\text{O} \]

water
• Bicarbonate ion (HCO$_3^-$) is an important base in the body and is especially abundant in the blood.
• Ammonia (NH$_3$), a natural waste product of protein breakdown, is also a base
• Ammonia has a pair of unshared electrons that strongly attract protons
• When accepting a proton, ammonia becomes an ammonium ion:

$$\text{NH}_3 + \text{H}^+ \rightarrow \text{NH}_4^+$$

ammonium ion
STRONG AND WEAK BASES
Strong Bases

• Remember, bases are proton acceptors
• Strong bases dissociate easily in water and quickly tie up H⁺
  – Hydroxides
• In contrast, weak bases:
  – Sodium bicarbonate or baking soda
  – Dissociate incompletely and reversibly and are slower to accept protons

• Because sodium bicarbonate accepts a relatively small amount of protons
  – Its released bicarbonate ion is described as a weak base
pH: Acid-Base Concentration

• As the concentration of hydrogen ions in a solution increase:
  – The more acidic the solution becomes
• As the level of hydroxide ions increases:
  – The more basic, or alkaline, the solution becomes
pH: Acid-Base Concentration

- Clinically, the concentration of hydrogen ions in the body is measured in units called *pH units*.
- pH scale runs from 0 to 14 and is logarithmic.
• Each successive unit change in pH represents a tenfold change in hydrogen ion concentration.
pH of a solution, therefore, is defined as the negative logarithm, to the base 10, of the hydrogen ion concentration \([H^+]\) in moles per liter, or \(-\log H^+\):

\[
pH = -\log_{10} [H^+]
\]
pH: Acid-Base Concentration

• When the pH is 7 ($H^+ = 10^{-7}$ mol/liter)
  – Number of hydrogen ions precisely equals the number of hydroxide ions ($OH^-$)

• And the solution is neutral
  – Neither acidic or basic
pH: Acid-Base Concentration

• Pure water has a neutral pH of 7, or $10^{-7}$ mol/liter (0.0000001 mol/liter) of hydrogen ions.

• A solution with a pH below 7, is acidic
  – There are more hydrogen ions than hydroxide ions
pH: Acid-Base Concentration

• For example, a solution with a pH of 6 has 10 times more hydrogen ions than a solution with a pH of 7
pH: Acid-Base Concentration

• A solution with a pH greater than 7, is alkaline
  – Hydroxide ions outnumber the hydrogen ions
• For example, a solution with a pH of 8 has 10 times more hydroxide ions than a solution with a pH of 7
pH: Acid-Base Concentration

• Thus, as the hydrogen ion concentration increases
  – Hydroxide ion concentration falls, and vice versa.
The pH scale represents the number of hydrogen ions in a substance. The concentration of hydrogen ions ($\text{H}^+$) and the corresponding hydroxyl concentration ($\text{OH}^-$) for each representative substance is also provided. Note that when the pH is 7.0, the amount of $\text{H}^+$ and $\text{OH}^-$ are equal and the solution is neutral.
Chemical Buffer Systems and Acid-Base Balance

- Chemical buffers resist pH changes and are the body’s first line of defense.
- Ability of an acid-base mixture to resist sudden changes in pH is called its buffer action.
Chemical Buffer Systems and Acid-Base Balance

• Tissue cells and vital organs of the body are extremely sensitive to even the slightest change in the pH environment
In high concentrations, both acids and bases can be extremely damaging to living cells

– Essentially every biological process within the body is disrupted
• Buffers work against sudden and large changes in the pH of body fluids by
  1. Releasing hydrogen ions (acting as acids) when the pH increases, and
  2. Binding hydrogen ions (acting as bases) when the pH decreases.
• Three major chemical buffer systems in the body are the:
  – Carbonic acid-bicarbonate buffer system
  – Phosphate buffer system
  – Protein buffer system
• The carbonic acid-bicarbonate buffer system
  – Plays an extremely important role in maintaining pH homeostasis of the blood.
Carbonic acid ($H_2CO_3$) dissociates reversibly and releases bicarbonate ions ($HCO_3^-$) and protons ($H^+$) as follows:

**Response to an increase in pH**

\[
H_2CO_3 \rightleftharpoons HCO_3^- + H^+
\]

$H^+$ donor (weak acid) \quad Response to a decrease in pH \quad H^+ acceptor (weak proton)
• Under normal conditions, the ratio between the $\text{HCO}_3^-$ and $\text{H}_2\text{CO}_3$ in the blood is 20:1
  – See Figure 7-1
• Chemical equilibrium between carbonic acid (weak acid) and bicarbonate ion (weak base) works to resist sudden changes in blood pH.
How $\text{CO}_2$ Is Converted to $\text{HCO}_3^-$

Fig. 7-1. How $\text{CO}_2$ is converted to $\text{HCO}_3^-$ at the tissue sites. Most of the $\text{CO}_2$ that is produced at the tissue cells is carried to the lungs in the form of $\text{HCO}_3^-$. 

$$H_2\text{CO}_3 = P_{\text{CO}_2} \times 0.0301$$
• For example, when the blood pH increases (i.e., becomes more alkaline from the addition of a strong base), the equilibrium shifts to the right.
• A right shift forces more carbonic acid to dissociate, which in turn causes the pH to decrease.
• In contrast, when the blood pH decreases (i.e., becomes more acidic from the addition of a strong acid), the equilibrium moves to the left.

• A left shift forces more bicarbonate to bind with protons.
Carbonic Acid-Bicarbonate Buffer System and Acid-Base Balance

• Carbonic acid-bicarbonate buffer system converts:
  1. Strong bases to a weak base (bicarbonate ion), and
  2. Strong acids to a weak acid (carbonic acid)
• Blood pH changes are much less than they would be if this buffering system did not exist.
The Henderson-Hasselbalch Equation (H-H)

• H-H equation mathematically illustrates how the pH of a solution is influenced by the HCO$_3^-$ to H$_2$CO$_3$ ratio
  – The base to acid ratio
The Henderson-Hasselbalch Equation (H-H)

- H-H equation is written as follows:

\[
pH = pK + \log \frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]} \text{ (base)}
\]

- \(pK\) is derived from the dissociation constant of the acid portion of the buffer combination

- \(pK\) is 6:1 and, under normal conditions, the \(\text{HCO}_3^-\) to \(\text{H}_2\text{CO}_3\) ratio is 20:1
The Henderson-Hasselbalch Equation (H-H)

- Clinically, the dissolved CO$_2$ (PCO$_2$ x 0.03) can be used for the denominator of the H-H equations, instead of the H$_2$CO$_3$
- This is possible since the dissolved carbon dioxide is in equilibrium with, and directly proportional to, the blood [H$_2$CO$_3$]
Thus, the H-H equation can be written as follows:

\[ \text{pH} = \text{pK} + \log \frac{[\text{HCO}_3^-]}{[\text{PCO}_2 \times 0.03]} \]
• When the $\text{HCO}_3^-$ is 24 mEq/L, and the $\text{PaCO}_2$ is 40 mm Hg, the base to acid ratio is 20:1 and the pH is 7.4 (normal).

• H-H equation confirms the 20:1 ratio and pH of 7.4 as follows:
H-H Equation Applied During Normal Conditions

\[ \text{pH} = \text{pK} + \log \frac{[\text{HCO}_3^-]}{[\text{PCO}_2 \times 0.03]} \]

\[ = 6.1 + \log \frac{24 \text{ mEq/L}}{(40 \times 0.03)} \]

\[ = 6.1 + \log \frac{24 \text{ mEq/L}}{(1.2 \text{ mEq/L})} \]

\[ = 6.1 + \log 20 \quad (20:1 \text{ ratio}) \]

\[ = 6.1 + 1.3 \]

\[ = 7.4 \]
H-H Equation Applied During Abnormal Conditions

• When the HCO$_3^-$ is 29 mEq/L, and the PaCO$_2$ is 80 mm Hg, the base to acid ratio decreases to 12:1 and the pH is 7.18 (acidic)

• H-H equation confirms the 12:1 ratio and the pH of 7.18 as follows:
H-H Equation Applied During Abnormal Conditions

\[
pH = pK + \log \frac{[\text{HCO}_3^-]}{[\text{PCO}_2 \times 0.03]}
\]

\[
= 6.1 + \log \frac{29 \text{ mEq/L}}{(80 \times 0.03)}
\]

\[
= 6.1 + \log \frac{29 \text{ mEq/L}}{(2.4 \text{ mEq/L})}
\]

\[
= 6.1 + \log 12 \quad (12:1 \text{ ratio})
\]

\[
= 6.1 + 1.08
\]

\[
= 7.18
\]
H-H Equation Applied During Abnormal Conditions

• In contrast, when the HCO₃⁻ is 20 mEq/L, and the PaCO₂ is 20 mm Hg, the base to acid ratio increases to 33:1 and the pH is 7.62 (alkalotic)

• H-H equation confirms the 33:1 ratio and the pH of 7.62 as follows:
H-H Equation Applied During Abnormal Conditions

\[ pH = pK + \log \left( \frac{[HCO_3^-]}{[PCO_2 \times 0.03]} \right) \]

\[ = 6.1 + \log \left( \frac{20 \text{ mEq/L}}{(20 \times 0.03)} \right) \]

\[ = 6.1 + \log \left( \frac{20 \text{ mEq/L}}{0.6 \text{ mEq/L}} \right) \]

\[ = 6.1 + \log \left( \frac{33}{1} \right) \quad (33:1 \text{ ratio}) \]

\[ = 6.1 + 1.52 \]

\[ = 7.62 \]
• Clinically, the H-H equation can be used to calculate the pH, $[\text{HCO}_3^-]$, or PCO$_2$ when any two of these three variables are known. $[\text{HCO}_3^-]$ is solved as follows:

$$[\text{HCO}_3^-] = \text{antilog} \ (7.40 - 6.1) \times (\text{PCO}_2 \times 0.03)$$

• PCO$_2$ is determined as follows:

$$\text{PCO}_2 = \frac{[\text{HCO}_3^-]}{(\text{antilog} \ [\text{pH} - 6.1] \times 0.03)}$$
Clinical Application of H-H Equation

• H-H equation may be helpful in cross checking the validity of the blood gas reports when the pH, PCO$_2$, and [HCO$_3^-$] values appear out of line.

• It may also be useful in estimating what changes to expect when any one of the H-H equation components is altered.
• For example, consider the case example that follows:
Case

• A mechanically ventilated patient has a pH of 7.54, a PaCO$_2$ of 26 mm Hg, and a HCO$_3^-$ of 22 mEq/L.
• The physician asks the respiratory practitioner to adjust the patient’s PaCO$_2$ to a level that will decrease the pH to 7.45.
• Using the H-H equation, the PaCO\textsubscript{2} change needed to decrease the pH to 7.45 can be estimated as follows:
\[
\text{PCO}_2 = \frac{[\text{HCO}_3^-]}{(\text{antilog} \ [\text{pH} - 6.1] \times 0.03)} \\
= \frac{22}{\text{antilog} \ (7.45 - 6.1) \times 0.03}
\]
\[
\frac{22}{\text{antilog } (1.35) \times 0.03}
\]

\[
= \frac{22}{22.38 \times 0.03}
\]

\[
= \frac{22}{0.67}
\]

\[
= 32.8 \text{ or } 33 \text{ mm Hg}
\]
• Thus, increasing the PaCO$_2$ to about 33 mm Hg should move the patient’s pH level close to 7.45.

• In this case, the respiratory practitioner would begin by either decreasing the tidal volume, or the respiratory rate, on the mechanical ventilator.
Clinical Application of H-H Equation

• After the ventilator changes are made
  – Another arterial blood gas should be obtained in about 20 minutes

• pH and PaCO₂ should be reevaluated
  – Followed by appropriate ventilator adjustments if necessary
Phosphate Buffer System and Acid-Base Balance

• Function of the phosphate buffer system is almost identical to that of the carbonic acid-bicarbonate buffer system
• Primary components of the phosphate buffer system are the:
  – Sodium salts of dihydrogen phosphate ($H_2PO_4^-$), and
  – Monohydrogen phosphate ($HPO_4^{2-}$)
Phosphate Buffer System and Acid-Base Balance

- NaH$_2$PO$_4$ is a weak acid
- Na$_2$HPO$_4$, which has one less hydrogen atom, is a weak base
Phosphate Buffer System and Acid-Base Balance

• When H\(^+\) ions are released by a strong acid, the phosphate buffer system works to inactivate the acidic effects of the H\(^+\) as follows:

\[
\text{HCl} + \text{Na}_2\text{HPO}_4 \xrightarrow{\Delta} \text{NaH}_2\text{PO}_4 + \text{NaCl}
\]

strong acid \hspace{1cm} \text{weak base} \hspace{1cm} \text{weak acid} \hspace{1cm} \text{salt}
Phosphate Buffer System and Acid-Base Balance

• On the other hand, strong bases are converted to weak bases as follows:

\[
\text{NaOH} + \text{NaH}_2\text{PO}_4 \xrightarrow{\Delta} \text{Na}_2\text{HPO}_4 + \text{H}_2\text{O}
\]

strong base weak acid weak base water
Phosphate Buffer System and Acid-Base Balance

• Phosphate buffer system is not an effective buffer for blood plasma
• It is an effective buffer system in urine and in intracellular fluid where the phosphate levels are typically greater.
Phosphate Buffer System and Acid-Base Balance

• Body’s most abundant and influential supply of buffers is the protein buffer system
• Its buffers are found in the proteins in the plasma and cells
Phosphate Buffer System and Acid-Base Balance

• In fact, about 75 percent of the buffering power of body fluids is found in the intracellular proteins.
Proteins are polymers of amino acids

Some have exposed groups of atoms known as organic acid (carboxyl) groups (—COOH), which dissociate and liberate H⁺ in response to a rising pH:

\[ R^* -\text{COOH} \rightarrow R -\text{COO}^- + H^+ \]

*R represents the entire organic molecule, which is composed of many atoms
• In contrast, other amino acids consist of exposed groups that can function as bases and accept $H^+$.

• For example, an exposed $-\text{NH}_2$ group can bond with hydrogen ions to form $-\text{NH}_3^+$:

$$R\ -\text{NH}_2 + H^+ \Delta R\ -\text{NH}_3^+$$
• Because this reaction ties up free hydrogen ions, it prevents the solution from becoming too acidic.

• In addition, a single protein molecule can function as either an acid or a base relative to its pH environment.
Phosphate Buffer System and Acid-Base Balance

• Protein molecules that have a reversible ability are called amphoteric molecules
Phosphate Buffer System and Acid-Base Balance

• Hemoglobin in red blood cells is a good example of a protein the works as an intracellular buffer.

• As discussed earlier, CO\textsubscript{2} released at the tissue cells quickly forms H\textsubscript{2}CO\textsubscript{3}, and then dissociates into H\textsuperscript{+} and HCO\textsubscript{3}\textsuperscript{−} ions.

• See Figure 7-1.
How CO₂ Is Converted to HCO₃⁻

Fig. 7-1. How CO₂ is converted to HCO₃⁻ at the tissue sites. Most of the CO₂ that is produced at the tissue cells is carried to the lungs in the form of HCO₃⁻.
• At the same time, the hemoglobin is unloading oxygen at the tissue sites and becoming reduced hemoglobin.

• Because reduced hemoglobin carries a negative charge
  – Free $H^+$ ions quickly bond to the hemoglobin anions
Phosphate Buffer System and Acid-Base Balance

• This action reduces the acidic effects of the \( \text{H}^+ \) on the pH

• In essence, the \( \text{H}_2\text{CO}_3 \), which is a weak acid, is buffered by an even weaker acid—the hemoglobin protein
The Respiratory System and Acid-Base Balance

- Respiratory system does not respond as fast as the chemical buffer systems.
- However, it has up to two times the buffering power of all of the chemical buffer systems combined.
• CO₂ produced by the tissue cells enters the red blood cells and is converted to HCO₃⁻ ions as follows:

\[ \text{CO}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{CO}_3 \rightleftharpoons \text{H}^+ + \text{HCO}_3^- \]
• The first set of double arrows illustrates a reversible equilibrium between the dissolved carbon dioxide and the water on the left
  – And carbonic acid on the right
The second set of arrows shows a reversible equilibrium between carbonic acid on the left and hydrogen and bicarbonate ions on the right.
• Because of this relationship, an increase in any of these chemicals causes a shift in the opposite direction

• Note also that the right side of this equation is the same as that for the carbonic acid-bicarbonate buffer system
• Under normal conditions, the volume of CO₂ eliminated at the lungs is equal to the amount of CO₂ produced at the tissues.
• When the CO$_2$ is unloaded at the lungs, the preceding equation flows to the left, and causes the H$^+$ generated from the carbonic acid to transform back to water.
  – See Figure 7-2
Fig. 7-2. How $\text{HCO}_3^-$ is transformed back into $\text{CO}_2$ and eliminated in the alveoli.
• Because of the protein buffer system, the $H^+$ generated by the $CO_2$ transport system is not permitted to increase
  – Therefore, it has little or no effect on blood pH
• Under abnormal conditions, the respiratory system quickly responds by either increasing or decreasing the rate and depth of breathing
  – To compensate for acidosis or alkalosis, respectively
• For example, when the pH declines (e.g., metabolic acidosis caused by lactic acids)
  – Respiratory system responds by increasing the breathing depth and rate
• This action causes more CO$_2$ to be eliminated from the lungs and, therefore, pushes the preceding reaction to the left and reduces the H$^+$ concentration.

• This process works to return the acidic pH back to normal.
• On the other hand, when the pH rises:
  – Metabolic alkalosis caused by hypokalemia
  – Respiratory system responds by decreasing the breathing depth and rate
• This action causes less CO$_2$ to be eliminated from the lungs and, thus, moves the preceding reaction to the right and increases the H$^+$ concentration.

• This works to pull the alkalotic pH back to normal.
• Note: When the respiratory system is impaired for any reason, a serious acid-base imbalance can develop.
• For example:
  – Severe head trauma can cause a dramatic increase in the depth and rate of breathing that is completely unrelated to the CO$_2$ concentration.
  – When this happens, the volume of CO$_2$ expelled from the lungs will be greater than amount of CO$_2$ produced at the tissue cells.
• In other words, hyperventilation is present.
• This condition causes the pH to increase and respiratory alkalosis is said to exist.
In contrast:

- The ingestion of barbiturates can cause a dramatic decrease in the depth and rate of breathing.
- When this occurs, the volume of CO$_2$ eliminated from the lungs is less than the amount of CO$_2$ produced at the tissue cells.
The Renal System and Acid-Base Balance

• Even though the chemical buffer systems can inactivate excess acids and bases momentarily, they are unable to eliminate them from the body.
Similarly, although the respiratory system can expel the volatile carbonic acid by eliminating CO$_2$, it cannot expel other acids generated by cellular metabolism.
• Only the renal system can rid the body of acids such as phosphoric acids, uric acids, lactic acids, and ketone acids
  – Also called fixed acids
• Only the renal system can regulate alkaline substances in the blood and restore chemical buffers that are used in managing H\(^+\) levels in extracellular fluids
  – Some HCO\(_3^-\), which helps to adjust H\(^+\) concentrations, is lost from the body when CO\(_2\) is expelled from the lungs.
When the extracellular fluids become acidic, the renal system retains $\text{HCO}_3^-$ and excretes $\text{H}^+$ ions into the urine

– This causes the blood pH to increase.
• When the extracellular fluids become alkaline, the renal system retains H\(^+\) and excretes basic substances primarily \(\text{HCO}_3^-\) into the urine
  – This causes the blood pH to decrease
THE ROLE OF THE 
$\text{PCO}_2/\text{HCO}_3^-/\text{pH}$ 
RELATIONSHIP IN 
ACID-BASE BALANCE
Acid-Base Balance Disturbances

- Normal bicarbonate (HCO$_3^-$) to carbonic acid (H$_2$CO$_3$) ratio in the blood plasma is 20:1.
- In other words, for every H$_2$CO$_3$ ion produced in blood plasma, 20 HCO$_3^-$ ions must be formed to maintain a 20:1 ratio (normal pH).
• Or, for every $\text{H}_2\text{CO}_3$ ion loss in the blood plasma, 20 $\text{HCO}_3^-$ ions must be eliminated to maintain a normal pH.

• In other words, the $\text{H}_2\text{CO}_3$ ion is 20 times more powerful than the $\text{HCO}_3^-$ ion in changing the blood pH.
• Under normal conditions, the 20:1 acid-base balance in the body is automatically regulated by the:
  – Chemical buffer systems
  – Respiratory system
  – Renal system
• However, these normal acid-base regulating systems have their limits.
• The bottom line is this:
  – The body’s normal acid-base watchdog systems cannot adequately respond to sudden changes in $H^+$ and $HCO_3^-$ concentrations
  • Regardless of the cause
For example:

- Hypoventilation causes the partial pressure of the alveolar carbon dioxide (PACO$_2$) to increase, which in turn causes the plasma PCO$_2$, HCO$_3^-$, and H$_2$CO$_3$ to all increase.
- This causes HCO$_3^-$ to H$_2$CO$_3$ ratio to decrease, and the pH to fall.
- See Figure 7-8
Alveolar hypoventilation causes the $PACO_2$ and the plasma $PCO_2$, $H_2CO_3$, and $HCO_3^-$ to increase. This action decreases the $HCO_3^-/H_2CO_3$ ratio, which in turn decreases the blood pH.
• Or, when the PACO$_2$ decreases, as a result of alveolar hyperventilation, the plasma PCO$_2$, HCO$_3^-$ and H$_2$CO$_3$ all decrease which in turn causes:
  – HCO$_3^-$ to H$_2$CO$_3$ ratio to increase, and the pH to rise
  – See Figure 7-9
Alveolar Hyperventilation

Fig. 7-9. Alveolar hyperventilation causes the PACO$_2$ and the plasma PCO$_2$, H$_2$CO$_3$, and HCO$_3^-$ to decrease. This action increases the HCO$_3^-$/H$_2$CO$_3$ ratio, which in turn increases the blood pH.
• Relationship between acute PCO$_2$ changes, and the resultant pH and HCO$_3^-$ changes that occur is graphically illustrated in the PCO$_2$/HCO$_3^-$/pH nomogram
  – See Figure 7-10
Figure 7-10  Nomogram of PCO$_2$/HCO$_3^-$/pH relationship.
Acid-Base Balance Disturbances

• PCO$_2$/HCO$_3^-$/pH nomogram is an excellent clinical tool that can be used to identify a specific acid-base disturbance

• Table 7-2 provides an overview of the common acid-base balance disturbances that can be identified on the PCO$_2$/ HCO$_3^-$/pH nomogram
Respiratory Acid-Base Disturbances

- Acute ventilatory failure (respiratory acidosis)
- Acute ventilatory failure with partial renal compensation
- Chronic ventilatory failure with complete renal compensation
- Acute alveolar hyperventilation (respiratory alkalosis)
- Acute alveolar hyperventilation with partial renal compensation
- Chronic alveolar hyperventilation with complete renal compensation
Metabolic Acid-Base Disturbances

- **Metabolic acidosis**
- Metabolic acidosis with partial respiratory compensation
- Metabolic acidosis with complete respiratory compensation
- Both metabolic and respiratory acidosis

- **Metabolic alkalosis**
- Metabolic alkalosis with partial respiratory compensation
- Metabolic alkalosis with complete respiratory compensation
- Both metabolic and respiratory alkalosis
The following sections will describe:

- The common acid-base disturbances, and
- How to identify them on the PCO$_2$/HCO$_3^-$/$\text{pH}$ nomogram.
RESPIRATORY ACID-BASE DISTURBANCES
• During acute ventilatory failure, $\text{PACO}_2$ progressively increases

• This action simultaneously causes an increase in the blood $\text{PCO}_2$, $\text{H}_2\text{CO}_3$ and $\text{HCO}_3^-$ levels
  – Causes blood pH to decrease or become acidic
  – See Figure 7-8
Fig. 7-8. Alveolar hypoventilation causes the PACO₂, H₂CO₃, and HCO₃⁻ to increase. This action decreases the HCO₃⁻/H₂CO₃ ratio, which in turn decreases the blood pH.
• Resultant pH and HCO$_3^-$ changes can be easily identified by using the left side of the red colored normal PCO$_2$ blood buffer bar located on the PCO$_2$/HCO$_3^-$/pH nomogram
  – Titled RESPIRATORY ACIDOSIS
  – See Figure 7-11
Acute ventilatory failure is confirmed when the reported PCO$_2$, pH and HCO$_3^-$ values all intersect within the red colored RESPIRATORY ACIDOSIS bar. For example, when the reported PCO$_2$ is 80 mm Hg, at a time when the pH is 7.18 and the HCO$_3^-$ is 28 mEq/L, acute ventilatory failure is confirmed.
Common Causes of Acute Ventilatory Failure

- **Chronic obstructive pulmonary disorders**
  - Pulmonary disorders such as chronic emphysema and chronic bronchitis can lead to acute ventilatory failure

- **Drug overdose**
  - such as narcotics or barbiturates can depress ventilation

- **General anesthesia**
  - Generally, anesthetics cause ventilatory failure
Common Causes of Acute Ventilatory Failure

• Head trauma
  – Severe trauma to the brain can cause acute ventilatory failure

• Neurologic disorders
  – Neurologic disorders such as Guillain-Barré Syndrome and Myasthenia Gravis can lead to acute ventilatory failure.
Renal Compensation

• In the patient who hypoventilates for a long period of time, the kidneys will work to correct the decreased pH by retaining HCO$_3^-$ in the blood.
  – Chronic obstructive pulmonary disease
The presence of renal compensation is verified when the reported $\text{PCO}_2$, $\text{HCO}_3^-$, and pH values all intersect in the purple colored area shown in the upper left-hand corner of the $\text{PCO}_2 / \text{HCO}_3^- / \text{pH}$ nomogram

- See Figure 7-12
Renal Compensation

Figure 7-12  Acute ventilatory failure with partial renal compensation (also called partially compensated respiratory acidosis) is present when the reported pH and $\text{HCO}_3^-$ are both above the normal red colored $\text{PCO}_2$ blood buffer bar (in the purple colored area), but the pH is still less than normal. For example, when the $\text{PCO}_2$ is 80 mm Hg, at a time when the pH is 7.30 and the $\text{HCO}_3^-$ is 37 mEq/L, ventilatory failure with partial renal compensation is confirmed.
Acute ventilatory failure with partial renal compensation is present when:

- Reported pH and $\text{HCO}_3^-$ are both above normal red colored $\text{PCO}_2$ blood buffer bar (in the purple colored area), but pH is still less than normal
- See Figure 7-12
Acute Ventilatory Failure with Partial Renal Compensation

Figure 7-12  Acute ventilatory failure with partial renal compensation (also called partially compensated respiratory acidosis) is present when the reported pH and HCO₃⁻ are both above the normal red colored PCO₂ blood buffer bar (in the purple colored area), but the pH is still less than normal. For example, when the PCO₂ is 80 mm Hg, at a time when the pH is 7.30 and the HCO₃⁻ is 37 mEq/L, ventilatory failure with partial renal compensation is confirmed.
Chronic Ventilatory Failure with Complete Renal Compensation (Compensated Respiratory Acidosis)

- Present when the $\text{HCO}_3^-$ increases enough to cause the acidic pH to move back into the normal range
  - Above 42 mEq/L
  - See Figure 7-12
Acute ventilatory failure with partial renal compensation (also called partially compensated respiratory acidosis) is present when the reported pH and HCO$_3^-$ are both above the normal red colored PCO$_2$ blood buffer bar (in the purple colored area), but the pH is still less than normal. For example, when the PCO$_2$ is 80 mm Hg, at a time when the pH is 7.30 and the HCO$_3^-$ is 37 mEq/L, ventilatory failure with partial renal compensation is confirmed.
• During acute alveolar hyperventilation, the PACO$_2$ will decrease and allow more CO$_2$ molecules to leave the pulmonary blood.
  – Hyperventilation due to pain and/or anxiety
Acute Alveolar Hyperventilation (Respiratory Alkalosis)

- This action simultaneously causes a decrease in the blood $\text{PCO}_2$, $\text{H}_2\text{CO}_3$, and $\text{HCO}_3^-$ levels
  - Which, in turn, causes the blood pH to increase, or become more alkaline
  - See Figure 7-9
Alveolar Hyperventilation

Fig. 7-9. Alveolar hyperventilation causes the \( P_{\text{ACO}_2} \) and the plasma \( \text{PCO}_2, \text{H}_2\text{CO}_3, \) and \( \text{HCO}_3^- \) to decrease. This action increases the \( \text{HCO}_3^- / \text{H}_2\text{CO}_3 \) ratio, which in turn increases the blood pH.
• Resultant pH and HCO$_3^-$ changes caused by an acute decrease in the PCO$_2$ level can be identified by using the right side of the red colored normal PCO$_2$ blood buffer bar located on the PCO$_2$/HCO$_3^-$/pH nomogram
  – Titled RESPIRATORY ALKALOSIS
  – See Figure 7-13
Acute Alveolar Hyperventilation

Figure 7-13. Acute alveolar hyperventilation is confirmed when the reported PCO₂, pH and HCO₃⁻ values all intersect within the red colored RESPIRATORY ALKALOSIS bar. For example, when the reported PCO₂ is 25 mm Hg, at a time when the pH is 7.55 and the HCO₃⁻ is 21 mEq/L, acute alveolar hyperventilation is confirmed.
Common Causes of Acute Ventilatory Failure

- **Hypoxia**
  - Any cause of hypoxia (e.g., lung disorders, high altitudes, and heart disease) can cause acute alveolar hyperventilation.

- **Pain, anxiety, and fever**
  - Relative to the degree of pain, anxiety, and fever, hyperventilation may be seen.
Common Causes of Acute Ventilatory Failure

• Brain inflammation
  – Relative to the degree of cerebral inflammation, hyperventilation may be seen.

• Stimulant drugs
  – Agents such as amphetamines can cause alveolar hyperventilation.
Renal Compensation

• In the patient who hyperventilates for a long period of time, the kidneys will work to correct the increased pH by excreting excess $\text{HCO}_3^-$ in the urine.
  – A patient who has been overly mechanically hyperventilated for more than 24 to 48 hours
Renal Compensation

• The presence of renal compensation is verified when the reported PCO$_2$, HCO$_3^-$, and pH values all intersect in the green colored area shown in the lower right-hand corner of the PCO$_2$/HCO$_3^-$/pH nomogram
  – See Figure 7-14
Alveolar Hyperventilation (with Partial Renal Compensation)

Figure 7-14 Alveolar hyperventilation with partial renal compensation (also called partially compensated respiratory alkalosis) is present when the reported pH and $\text{HCO}_3^-$ are both below the normal red colored PCO$_2$ blood buffer bar (in the green colored area), but the pH is still greater than normal. For example, when the PCO$_2$ is 20 mm Hg, at a time when the pH is 7.50 and the $\text{HCO}_3^-$ is 15 mEq/L, alveolar hyperventilation with partial renal compensation is confirmed.
• Alveolar hyperventilation with partial renal compensation is present when:
  – Reported pH and $\text{HCO}_3^-$ are both below the normal red colored PCO$_2$ blood buffer bar (in the green colored area), but the pH is still greater than normal.
  – See Figure 7-14
Alveolar hyperventilation with partial renal compensation (also called partially compensated respiratory alkalosis) is present when the reported pH and HCO₃⁻ are both below the normal red colored PCO₂ blood buffer bar (in the green colored area), but the pH is still greater than normal. For example, when the PCO₂ is 20 mm Hg, at a time when the pH is 7.50 and the HCO₃⁻ is 15 mEq/L, alveolar hyperventilation with partial renal compensation is confirmed.
Chronic alveolar hyperventilation with complete renal compensation is present when the $\text{HCO}_3^-$ level decreases enough to return the alkalotic pH to normal.

- Which, in this case, would be below 14 mEq/L
- See Figure 7–14
Chronic Alveolar Hyperventilation with Complete Renal Compensation (Compensated Respiratory Alkalosis)

Figure 7-14  Alveolar hyperventilation with partial renal compensation (also called partially compensated respiratory alkalosis) is present when the reported pH and HCO$_3^-$ are both below the normal red colored PCO$_2$ blood buffer bar (in the green colored area), but the pH is still greater than normal. For example, when the PCO$_2$ is 20 mm Hg, at a time when the pH is 7.50 and the HCO$_3^-$ is 15 mEq/L, alveolar hyperventilation with partial renal compensation is confirmed.
General Comments

• As a general rule, the kidneys do not overcompensate for an abnormal pH.
• If the patient’s blood pH becomes acidic for a long period of time due to hypoventilation, the kidneys will not retain enough $\text{HCO}_3^-$ for the pH to climb higher than 7.40.
• The opposite is also true:
  – Should the blood pH become alkalotic for a long period of time due to hyperventilation, the kidneys will not excrete enough $\text{HCO}_3^-$ for the pH to fall below 7.40.
• However, there is one important exception to this rule:
  – In persons who chronically hypoventilate for a long period of time, it is not uncommon to find a pH greater than 7.40 (e.g., 7.43 or 7.44)
  – Patients with chronic emphysema or chronic bronchitis
• This is due to water and chloride ion shifts that occur between the intercellular and extracellular spaces when the renal system works to compensate for a decreased blood pH.

• This action causes an overall loss of blood chloride (hypochloremia)
  – Hypochloremia increases the blood pH
To Summarize

• The lungs play an important role in maintaining the $\text{PCO}_2$, $\text{HCO}_3^-$, and pH levels on a moment-to-moment basis.

• The kidneys, on the other hand, play an important role in balancing the $\text{HCO}_3^-$ and pH levels during long periods of hyperventilation or hypoventilation.
METABOLIC ACID-BASE IMBALANCES
Metabolic Acidosis

• The presence of other acids, not related to an increased PCO$_2$ level, can also be identified on the PCO$_2$/HCO$_3^-$/pH nomogram.

• Clinically, this condition is called metabolic acidosis.
• Metabolic acidosis is present when the PCO$_2$ reading is within the normal range (35 to 45 mm Hg)—but not within the red colored normal blood buffer line when compared to the reported HCO$_3^-$ and pH levels.
• This is because the pH and \( \text{HCO}_3^- \) readings are both lower than expected for a normal \( \text{PCO}_2 \) level.
Metabolic Acidosis

- When the reported pH and HCO$_3^-$ levels are both lower than expected for a normal PCO$_2$ level, the PCO$_2$ reading will drop into the purple colored bar—
  - Titled METABOLIC ACIDOSIS
Metabolic Acidosis

• In short, the pH, HCO$_3^-$, and PCO$_2$ readings will all intersect within the purple colored METABOLIC ACIDOSIS bar
  – See Figure 7-15
When the reported pH and HCO$_3^-$ levels are both lower than expected for a normal PCO$_2$ level, the PCO$_2$ reading will drop into the purple colored bar titled METABOLIC ACIDOSIS. For example, when the reported PCO$_2$ is 40 mm Hg (normal), at a time when the pH is 7.25 and the HCO$_3^-$ is 17 mEq/L, metabolic acidosis is confirmed.
Common Causes of Metabolic Acidosis

- Lactic acidosis
  - Fixed acids
- Ketoacidosis
  - Fixed acids
Common Causes of Metabolic Acidosis

- Salicylate intoxication
  - Aspirin overdose
  - Fixed acids
- Renal failure
- Uncontrolled diarrhea

Table 7-4
Anion Gap

• To determine if a patient’s metabolic acidosis is caused by either:
  1. The accumulation of fixed acids (e.g., lactic acids, keto acids, or salicylate intoxication), or
  2. By an excessive loss of $\text{HCO}_3^-$
• According to the law of electroneutrality:
  – The total number of plasma positively charged ions (cations) must equal the total number of plasma negatively charged ions (anions) in the body fluids.

• To determine the anion gap, the most commonly measured cations are sodium (Na\(^+\)) ions.
Anion Gap

• Most commonly measured anions are the chloride (Cl\textsuperscript{−}) ions and bicarbonate (HCO\textsubscript{3}\textsuperscript{−}) ions
• The normal plasma concentration of these cations and anions are as follows:
  \begin{align*}
    \text{Na}^+ & : \quad 140 \text{ mEq/L} \\
    \text{Cl}^- & : \quad 105 \text{ mEq/L} \\
    \text{HCO}_3^- & : \quad 24 \text{ mEq/L}
  \end{align*}
Anion Gap

• Mathematically, the anion gap is the calculated difference between the Na\(^+\) ions and the sum of the HCO\(_3^-\) and Cl\(^-\) ions:

\[
\text{Anion gap} = [\text{Na}^+] - ([\text{Cl}^-] + [\text{HCO}_3^-])
\]

\[
= 140 - 105 + 24
\]

\[
= 140 - 129
\]

\[
= 11 \text{ mEq/L}
\]
Anion Gap

• The normal anion gap range (or the range of the unmeasured ions) is 9 to 14 mEq/L.
• An anion gap greater than 14 mEq/L represents metabolic acidosis.
Anion Gap

• An elevated anion gap is most commonly caused by the accumulation of fixed acids
  – Lactic acids
  – Keto acids
  – Salicylate intoxication
Anion Gap

• This is because the H\(^+\) ions that are generated by the fixed acids chemically react with—and are buffered by—the plasma HCO\(_3^-\)
• This action causes:
  1. The HCO\(_3^-\) concentration to decrease, and
  2. The anion gap to increase
• Clinically, when the patient presents with both metabolic acidosis and an increased anion gap, the respiratory care practitioner must investigate further to determine the source of the fixed acids.
  – This needs to be done in order to appropriately treat the patient.
• For example, a metabolic acidosis caused:

1. By lactic acids justifies the need for oxygen therapy—to reverse the accumulation of the lactic acids, or
2. By ketone acids justifies the need for insulin—to reverse the accumulation of the ketone acids.
• Interestingly, metabolic acidosis caused by an excessive loss of \( \text{HCO}_3^- \) does not cause the anion gap to increase
  – Namely, renal disease or severe diarrhea
Anion Gap

• This is because as the $\text{HCO}_3^-$ concentration decreases, the $\text{Cl}^-$ concentration routinely increases to maintain electroneutrality.

• In other words, for each $\text{HCO}_3^-$ that is lost, a $\text{Cl}^-$ anion takes its place:
  – Namely, law of electroneutrality
Anion Gap

• This action maintains a normal anion gap
• Metabolic acidosis caused by a decreased $\text{HCO}_3^-$ is often called hyperchloremic metabolic acidosis
To Summarize

• When metabolic acidosis is accompanied by an increased anion gap, the most likely cause of the acidosis is fixed acids
  – Lactic acids
  – Ketoacids
  – Salicylate intoxication
To Summarize

• Or, when a metabolic acidosis is seen with a normal anion gap, the most likely cause of the acidosis is an excessive loss of $\text{HCO}_3^-$
  – Caused by renal disease or severe diarrhea
The immediate compensatory response to metabolic acidosis is an increased ventilatory rate.
– This action causes the PaCO$_2$ to decline
Metabolic Acidosis with Respiratory Compensation

• As the PCO$_2$ decreases, the H$^+$ concentration decreases
  – This action works to offset the metabolic acidosis
  – See Figure 7-9
Alveolar hyperventilation causes the PACO$_2$ and the plasma PCO$_2$, H$_2$CO$_3$, and HCO$_3^-$ to decrease. This action increases the HCO$_3^-$/H$_2$CO$_3$ ratio, which in turn increases the blood pH.
Metabolic Acidosis with Partial Respiratory Compensation

• When pH, HCO$_3^-$, and PCO$_2$ all intersect in the yellow colored area of the PCO$_2$/HCO$_3^-$/pH nomogram, metabolic acidosis with partial respiratory compensation is present.
• In other words, the PaCO₂ has decreased below the normal range, but the pH is still below normal
  – See Figure 7-16
Metabolic Acidosis with Partial Respiratory Compensation

Figure 7-16  When the pH, HCO₃⁻, and PCO₂ all intersect in the yellow colored area of the PCO₂/HCO₃⁻/pH nomogram, metabolic acidosis with partial respiratory compensation is present. For example, when the PCO₂ is 25 mm Hg, at a time when the pH is 7.30 and the HCO₃⁻ is 12 mEq/L, metabolic acidosis with partial respiratory compensation is confirmed.
• Metabolic acidosis with complete respiratory compensation is present when the PaCO$_2$ decreases enough to move the acidic pH back to the normal range
  – Which, in this case, would be below 20 mm Hg
  – See Figure 7-16
Both Metabolic and Respiratory Acidosis

- When the pH, HCO$_3^-$, and PCO$_2$ readings all intersect in the orange colored area of the PCO$_2$/HCO$_3^-$/pH nomogram:
  - Both metabolic and respiratory acidosis are present
  - See Figure 7-17
Both Metabolic and Respiratory Acidosis

Figure 7-17 When the pH, HCO₃⁻, and PCO₂ readings all intersect in the orange colored area of the PCO₂/HCO₃⁻/pH nomogram, both metabolic and respiratory acidosis are present. For example, if the reported PCO₂ is 70 mm Hg, at a time when the pH is 7.10 and the HCO₃⁻ is 21 mEq/L, both metabolic and respiratory acidosis are present.
Both Metabolic and Respiratory Acidosis

- Both metabolic and respiratory acidosis are commonly seen in patients with acute ventilatory failure
  - Causes blood $\text{PCO}_2$ to increase (respiratory acidosis), and
  - $\text{PO}_2$ to decrease (metabolic acidosis—caused by lactic acids)
Metabolic Alkalosis

• Presence of other bases, not related to a decreased PCO$_2$ level or renal activity, can also be identified on the PCO$_2$/HCO$_3^-$/pH nomogram.

• Clinically, this condition is called metabolic alkalosis
Metabolic Alkalosis

• Metabolic alkalosis is present when the PCO$_2$ reading is within the normal range (35 to 45 mm Hg)—but not within the red normal blood buffer line when compared to the reported pH and HCO$_3^-$ levels.

• This is because the pH and HCO$_3^-$ readings are both higher than expected for a normal PCO$_2$ level.
• When reported pH and $\text{HCO}_3^-$ levels are both higher than expected for a normal PCO$_2$ level, PCO$_2$ reading will move up into the purple colored bar titled METABOLIC ALKALOSIS
  – pH, $\text{HCO}_3^-$, and PCO$_2$ readings will all intersect within the purple colored bar.
  • See Figure 7-18
Metabolic Alkalosis

Figure 7-18. When the reported pH and HCO$_3^-$ levels are both higher than expected for a normal PCO$_2$ level, the PCO$_2$ reading will move up into the purple colored bar titled METABOLIC ALKALOSIS. For example, when the reported PCO$_2$ is 40 mm Hg (normal), at a time when the pH is 7.50 and the HCO$_3^-$ is 31 mEq/L, metabolic alkalosis is confirmed.
Common Causes of Metabolic Alkalosis

- Hypokalemia
- Hypochloremia
- Gastric suctioning or vomiting
- Excessive administration of corticosteroids
- Excessive administration of sodium bicarbonate
- Diuretic therapy
- Hypovolemia
Metabolic Alkalosis with Respiratory Compensation

• The immediate compensatory response to metabolic alkalosis is a decreased ventilatory rate.
  – This action causes the PaCO$_2$ to rise
Metabolic Alkalosis with Respiratory Compensation

• As the PCO$_2$ increases, the H$^+$ concentration increases
  – This action works to offset the metabolic alkalosis
  – See Figure 7-8
Metabolic Alkalosis with Respiratory Compensation

Fig. 7-8. Alveolar hypoventilation causes the PACO$_2$ and the plasma PCO$_2$, H$_2$CO$_3$, and HCO$_3^-$ to increase. This action decreases the HCO$_3^-$/H$_2$CO$_3$ ratio, which in turn decreases the blood pH.
Metabolic Alkalosis with Respiratory Compensation

• When pH, $\text{HCO}_3^-$, and $\text{PCO}_2$ all intersect in the pink colored area of the $\text{PCO}_2/ \text{HCO}_3^-/\text{pH}$ nomogram, metabolic alkalosis with partial respiratory compensation is present.
• In other words, the PaCO$_2$ has increased above the normal range, but the pH is still above normal
  – See Figure 7-19
Metabolic Alkalosis with Respiratory Compensation

Figure 7-19. When the pH, HCO$_3^-$, and PCO$_2$ all intersect in the pink colored area of the PCO$_2$/HCO$_3^-$/pH nomogram, metabolic alkalosis with partial respiratory compensation is present. For example, when the PCO$_2$ is 60 mm Hg, at a time when the pH is 7.50 and the HCO$_3^-$ is 46 mEq/L, metabolic alkalosis with partial respiratory compensation is present.
• Metabolic alkalosis with complete respiratory compensation is present when the PaCO$_2$ increases enough to move the alkalotic pH back to the normal range
  – Which, in this case, would be above 65 to 68 mm Hg
  – See Figure 7-19
Metabolic Alkalosis with Complete Respiratory Compensation

Figure 7-19 When the pH, HCO$_3^-$, and PCO$_2$ all intersect in the pink colored area of the PCO$_2$/HCO$_3^-$/pH nomogram, metabolic alkalosis with partial respiratory compensation is present. For example, when the PCO$_2$ is 60 mm Hg, at a time when the pH is 7.50 and the HCO$_3^-$ is 46 mEq/L, metabolic alkalosis with partial respiratory compensation is present.
Both Metabolic and Respiratory Alkalosis

• When the pH, HCO$_3^-$, and PCO$_2$ readings all intersect in the blue colored area of the PCO$_2$/HCO$_3^-$/pH nomogram,
  – Both metabolic and respiratory alkalosis are present
  – See Figure 7-20
Both Metabolic and Respiratory Alkalosis

Figure 7-20. When the pH, HCO$_3^-$, and PCO$_2$ readings all intersect in the blue colored area of the PCO$_2$/HCO$_3^-$/pH nomogram, both metabolic and respiratory alkalosis are present. For example, if the reported PCO$_2$ is 25 mm Hg, at a time when pH is 7.62 and the HCO$_3^-$ is 25 mEq/L, both metabolic and respiratory alkalosis are present.
Base Excess/Deficit

• The PCO$_2$/HCO$_3^-$/pH nomogram also serves as an excellent tool to calculate the patient’s total base excess/deficit.
• By knowing the base excess/deficit, non-respiratory acid-base imbalances can be quantified.
• The base excess/deficit is reported in milliequivalents per liter (mEq/L) of base above or below the normal buffer line of the $\text{PCO}_2/\text{HCO}_3^-/\text{pH}$ nomogram
• For example:

• If the pH is 7.25, and the $\text{HCO}_3^-$ is 17 mEq/L, at a time when the PaCO$_2$ is 40 mm Hg, the PCO$_2$/HCO$_3^-$/pH nomogram will confirm the presence of:
  – Metabolic acidosis, and
  – A base excess of -7 mEq/L
    • More properly called a base deficit of 7 mEq/L
  – See Figure 7-15
Metabolic Acidosis

Figure 7-15  When the reported pH and HCO$_3^-$ levels are both lower than expected for a normal PCO$_2$ level, the PCO$_2$ reading will drop into the purple colored bar titled METABOLIC ACIDOSIS. For example, when the reported PCO$_2$ is 40 mm Hg (normal), at a time when the pH is 7.25 and the HCO$_3^-$ is 17 mEq/L, metabolic acidosis is confirmed.
Base Excess/Deficit

- Metabolic acidosis may be treated by the careful intravenous infusion of sodium bicarbonate (NaHCO$_3$)
In contrast:

If the pH is 7.50, and the HCO$_3^-$ is 31 mEq/L, at a time when the PaCO$_2$ is 40 mm Hg, the PCO$_2$/HCO$_3^-$/pH nomogram will verify the presence of:

- Metabolic alkalosis, and
- A base excess of 7 mEq/L
- See Figure 7-18
Metabolic Alkalosis

Figure 7-18 When the reported pH and HCO$_3^-$ levels are higher lower than expected for a normal PCO$_2$ level, the PCO$_2$ reading will move up into the purple colored bar titled METABOLIC ALKALOSIS. For example, when the reported PCO$_2$ is 40 mm Hg (normal), at a time when the pH is 7.50 and the HCO$_3^-$ is 31 mEq/L, metabolic alkalosis is confirmed.
• Metabolic alkalosis is treated by:
  – Correcting underlying electrolyte problem
    • Namely, hypokalemia or hypochloremia
  – Administering ammonium chloride (NH₄Cl)
Example of Clinical Use of PCO$_2$/HCO$_3^-$/$p$H Nomogram

• It has been shown that the PCO$_2$/ HCO$_3^-$ /$p$H nomogram is an excellent clinical tool to confirm the presence of:
  – Respiratory acid-base imbalances,
  – Metabolic acid-base imbalances, or
  – A combination of a respiratory and metabolic acid-base imbalances
The clinical application cases at the end of this chapter further demonstrate the clinical usefulness of the PCO$_2$/HCO$_3^-$/$\text{pH}$ nomogram.
Clinical Application 1 Discussion

- How did this case illustrate …
  - How clinical signs and symptoms can sometimes be very misleading.
  - How the PCO$_2$/HCO$_3$/pH nomogram can be used to determine the cause of certain findings of arterial blood gas analysis.
Clinical Application 2 Discussion

• How did this case illustrate …
  – $\text{PCO}_2/\text{HCO}_3/\text{pH}$ nomogram can be used to confirm both a respiratory and metabolic acidosis.
  – $\text{PCO}_2/\text{HCO}_3/\text{pH}$ nomogram can be used to prevent the unnecessary administration of sodium bicarbonate during an emergency situation.