Blood Buffer Lesson Plan

IMPORTANT CONCEPTS TO REVIEW

1) Definitions of Acids and Bases (Brönsted-Lowry)
    a. *An acid is a proton (H^+) donor, a base is a proton acceptor*
2) Equilibrium
    a. LeChatelier's Principle
        i. *If a change is imposed on a system at equilibrium, the position of the equilibrium will shift in a direction that tends to reduce that change.*
    b. Equilibrium Expressions
        i. $K_a = [H_3O^+][A^-]/[HA] = [H^+][A^-]/[HA]$
        ii. $pKa = -\log[Ka]$
    c. pH
        i. $pH = -\log[H^+]$
    d. Henderson Hasselbalch Equation:
        i. $pH = pKa + \log([A^-]/[HA]) = pKa + \log([base]/[acid])$

DEFINITION OF A BUFFER – THE BASICS

A Buffer is a solution that:

1) Usually consists of a weak acid and a soluble salt, which contains the same anion as the weak acid. (**Note:** Can consist of a weak base and its soluble salt, which contains the same cation as the weak base)
    a. Example: Carbonic Acid/Bicarbonate Buffer System in Blood
        i. H_2CO_3 (weak acid, carbonic acid) + $NaHCO_3$ (soluble salt, sodium bicarbonate ion)
            1. **Note:** $NaHCO_3 \Leftrightarrow Na^+ + HCO_3^-$ (conjugate base of the weak acid, bicarbonate ion)
            2. **Note:** $H_2CO_3 \Leftrightarrow H^+ + HCO_3^-$ (conjugate base of the weak acid, bicarbonate ion)
        ii. Thus, anion in common is bicarbonate ion = HCO_3^-
2) Resists a change in pH, when either (basic) hydroxide ions (OH^-) or (acidic) protons (H^+ from H_3O^+) are added
    a. Example: Carbonic Acid/Bicarbonate Buffer System in Blood

i. **Remember**: H_2CO_3 (weak acid, carbonic acid) + $NaHCO_3$ (soluble salt, sodium bicarbonate ion)

ii. **Remember!** Carbonic acid is a weak acid, and thus remains in equilibrium with its conjugate base (Bicarbonate ion), unless the equilibrium is disturbed by the addition of acid or base

iii. If a strong acid is added to the buffer solution, then the salt neutralizes it to a weak acid

1. $[Na^+][HCO_3^-]$ (salt) + H_3O^+ (strong acid) → H_2CO_3 (weak acid) + Na^+ + H_2O

iv. If a strong base is added to the buffer solution, then the weak acid neutralizes it to become the salt

1. H_2CO_3 (weak acid) + OH^- (strong base)+ Na^+ → H_2O (water) + $[Na^+][HCO_3^-]$ (soluble salt)

BLOOD BUFFER SYSTEM – SOME FACTS AND SOME MATH

1) Blood has a pH around 7.4, usually between 7.35 and 7.45
    a. Acidosis is a serious condition where the pH of blood is below a pH of 7.35
    b. Alkalosis is a serious condition where the pH of blood is above a pH of 7.45
    c. Blood pHs below 6.9 or above 7.9 are usually fatal if they last for more than a short time
2) To keep the blood at a pH of around 7.4, several biological buffers exists in the blood
    a. **bicarbonate buffer system** (We will focus on this one!)
    b. phosphate buffer system
    c. protein buffer system
3) **BICARBONATE BUFFER SYSTEM**
    a. Most important buffer in the blood buffer system
    b. Based on respiratory rates
        i. Carbonic Acid is formed by a reaction of carbon dioxide and water
            1. CO_2 (g) ⇔ CO_2 (aq) + H_2O ⇔ H_2CO_3 (Carbonic acid)
    c. The optimal pH of the bicarbonate buffer system is: 6.4
        i. The optimal pH is the pKa for the acid in the system
        ii. pKa comes from the equilibrium expression for the acid
            1. $K_a = [H^+][A^-]/[HA] = [H^+][HCO_3^-]/[H_2CO_3] = 4.3 \times 10^{-7}$
            2. $pKa = -\log[K_a] = 6.4$

d. **Hmmm????** But the pH of blood is 7.4. And the optimal pH of the buffer is 6.4, which is 1 pH away from 7.4. How does that work?
    i. **Remember!** Henderson-Hasselbalch Equation
        1. $pH = pKa + \log([HCO_3^-]/[H_2CO^3])$
            a. $7.4 = 6.4 + \log([HCO_3^-]/[H_2CO^3])$
            b. $1 = \log([HCO_3^-]/[H_2CO^3])$
            c. $1 = \log([10]/[1])$
        2. **Thus**, to maintain a pH of 7.4, the equilibrium needs to be maintained at 10 units of bicarbonate ion (conjugate base) for every 1 unit of carbonic acid (weak acid)
    ii. **Why is this so?**
        1. Natural metabolism produces much more acids than bases, thus more bicarbonate ions (conjugate base) is needed in the blood to neutralize the incoming acids

HOW TO PREPARE FOR THE GAME?

1) Blood has a pH of 7.4 at the start of the game
    a. **Remember!** Henderson-Hasselbalch Equation
        i. $pH = pKa + \log([HCO_3^-]/[H_2CO^3])$
        ii. $7.4 = 6.4 + \log([HCO_3^-]/[H_2CO^3])$
        iii. $1 = \log([HCO_3^-]/[H_2CO^3])$
        iv. $1 = \log([10]/[1])$
    b. **Thus,** at the start of the game, there are 10 units of bicarbonate ions for every 1 unit of carbonic acid
2) **BUT** Acidosis occurs so the pH lowers naturally when the body releases organic acids into the bloodstream. You need to release bicarbonate ion units to counteract the released acids.
3) **If** the pH of blood changes to 6.9
    a. **Use** the Henderson-Hasselbalch Equation again
        i. $pH = pKa + \log([HCO_3^-]/[H_2CO^3])$
        ii. $6.9 = 6.4 + \log([HCO_3^-]/[H_2CO^3])$
        iii. $0.5 = \log([HCO_3^-]/[H_2CO^3])$
        iv. $1 = \log([\sim 3]/[1])$
    b. **Thus,** you currently have 3 bicarbonate units to every 1 carbonic acid unit. To return to the equilibrium of 10 to 1, you need to **add 7 units** of bicarbonate

4) **If** the pH of blood changes to 6.4
    a. **Use** the Henderson-Hasselbalch Equation again
        i. $pH = pKa + \log([HCO_3^-]/[H_2CO^3])$
        ii. $6.4 = 6.4 + \log([HCO_3^-]/[H_2CO^3])$
        iii. $0 = \log([HCO_3^-]/[H_2CO^3])$
        $0 = \log([1]/[1])$
    b. **Thus,** you currently have 1 bicarbonate unit to every 1 carbonic acid unit. To return to the equilibrium of 10 to 1, you need to **add 9 units** of bicarbonate
5) **YOU** figure out the rest in between a pH change from 7.4 to 6.4 using the Excel worksheet.

HOW TO PLAY THE GAME?

1) **Click the hammer button**, **press the counter** (bottom left of screen) and **press the center tile** of the screen. A message will state the dead blood cell count of zero, the beginning pH and "Begin game."
2) **Set the speed button** to whatever rate you want. We suggest slow to start, but feel free to adjust as you wish.
3) **Click the start button**
4) **Press Space** (whenever you want) for a message stating what the pH is
    a. If the pH is 7.4, Equilibrium is achieved! (But it doesn't stay for long...)
    b. If the pH is above 7.4 you are experiencing alkalosis and you should wait for equilibrium to return (the body naturally releases acids into the blood)
    c. If the pH is below 7.4, you are experiencing acidosis (you need to respond to the acids release into the blood). You should add bicarbonate buffer units to the system to raise the pH. If you do nothing, pH will drop even lower.
        i. **Press a number button** to release the corresponding units of bicarbonate ions. (Example: The 2 Button releases 2 units of Bicarbonate
        ii. REMEMBER your excel worksheet!!! You already calculated the approximate amount of bicarbonate ions to release at each pH increment.
        iii. **Press Space** again to see if your pH returned to the correct equilibrium
    d. If the pH falls below 6.4 or above 7.9, you die!!! If you add too much bicarbonate buffer, you cause alkalosis and die. If you don't add enough bicarbonate buffer then the pH will continue to fall and you cause acidosis and die.
    e. If the dead red blood cells stack up and clog the artery, you die! Be careful!

Resources

Johll, M. E. **Investigating Chemistry: A Forensic Science Perspective**. W. H. Freeman Publishing: New York, NY. 2009

Zumdahl, S. S. **Chemistry**. D. C Heath Publishing: Lexington, MA. 1989

http://www.chemistry.wustl.edu/~edudev/LabTutorials/Buffer/Buffer.html

http://www.harpercollege.edu/tm-ps/chm/100/dgodambe/thedisk/bloodbuf/zback2.htm

http://www.brynmawr.edu/chemistry/Chem/Chem104lc/buffers.html

http://web.mnstate.edu/provost/Chem400_4.pdf

http://faculty.stcc.edu/AandP/AP/AP2pages/Units21to23/ph/buffers.htm