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**Test Bank for Lehninger  
Principles of Biochemistry 5th  
Edition by Nelson**

## Multiple Choice Questions

### 1. Weak interactions in aqueous systems

Page: 43–45      Difficulty: 2      Ans: D

Which of these statements about hydrogen bonds is *not* true?

- A) Hydrogen bonds account for the anomalously high boiling point of water.
- B) In liquid water, the average water molecule forms hydrogen bonds with three to four other water molecules.
- C) Individual hydrogen bonds are much weaker than covalent bonds.
- D) Individual hydrogen bonds in liquid water exist for many seconds and sometimes for minutes.
- E) The strength of a hydrogen bond depends on the linearity of the three atoms involved in the bond.

### 2. Weak interactions in aqueous systems

Page: 48      Difficulty: 2      Ans: A

A true statement about hydrophobic interactions is that they:

- A) are the driving force in the formation of micelles of amphipathic compounds in water.
- B) do not contribute to the structure of water-soluble proteins.
- C) have bonding energies of approximately 20–40 Kjoule per mole.
- D) involve the ability of water to denature proteins.
- E) primarily involve the effect of polar solutes on the entropy of aqueous systems.

### 3. Weak interactions in aqueous systems

Page: 48–49      Difficulty: 2      Ans: E

Hydrophobic interactions make important energetic contributions to:

- A) binding of a hormone to its receptor protein.
- B) enzyme-substrate interactions.
- C) membrane structure.
- D) three-dimensional folding of a polypeptide chain.
- E) all of the above are true.

### 4. Weak interactions in aqueous systems

Page: 51      Difficulty: 2      Ans: A

Dissolved solutes alter some physical (colligative) properties of the solvent water because they change the:

- A) concentration of the water.
- B) hydrogen bonding of the water.
- C) ionic bonding of the water.
- D) pH of the water.
- E) temperature of the water.

**5. Weak interactions in aqueous systems****Page: 51 Difficulty: 2 Ans: E**

Osmosis is movement of a:

- A) charged solute molecule (ion) across a membrane.
- B) gas molecule across a membrane.
- C) nonpolar solute molecule across a membrane.
- D) polar solute molecule across a membrane.
- E) water molecule across a membrane.

**6. Ionization of water, weak acids, and weak bases****Page: 54 Difficulty: 2 Ans: E**

A hydronium ion:

- A) has the structure  $\text{H}_3\text{O}^+$ .
- B) is a hydrated hydrogen ion.
- C) is a hydrated proton.
- D) is the usual form of one of the dissociation products of water in solution.
- E) all of the above are true.

**7. Ionization of water, weak acids, and weak bases****Page: 56 Difficulty: 2 Ans: A**

The pH of a solution of 1 M HCl is:

- A) 0
- B) 0.1
- C) 1
- D) 10
- E) -1

**8. Ionization of water, weak acids, and weak bases****Page: 56 Difficulty: 2 Ans: D**

The pH of a solution of 0.1 M NaOH is:

- A) 0.1
- B) 1.0
- C) 12.8
- D) 13
- E) 14

**9. Ionization of water, weak acids, and weak bases****Page: 56 Difficulty: 2 Ans: D**

Which of the following is true about the properties of aqueous solutions?

- A) A pH change from 5.0 to 6.0 reflects an increase in the hydroxide ion concentration ( $[\text{OH}^-]$ ) of 20%.
- B) A pH change from 8.0 to 6.0 reflects a decrease in the proton concentration ( $[\text{H}^+]$ ) by a factor of 100.
- C) Charged molecules are generally insoluble in water.
- D) Hydrogen bonds form readily in aqueous solutions.
- E) The pH can be calculated by adding 7 to the value of the pOH.

**10. Ionization of water, weak acids, and weak bases****Page: 56 Difficulty: 2 Ans: E**

The pH of a sample of blood is 7.4, while gastric juice is pH 1.4. The blood sample has:

- A) 0.189 times the  $[H^+]$  as the gastric juice.
- B) 5.29 times lower  $[H^+]$  than the gastric juice.
- C) 6 times lower  $[H^+]$  than the gastric juice.
- D) 6,000 times lower  $[H^+]$  than the gastric juice.
- E) a million times lower  $[H^+]$  than the gastric juice.

**11. Ionization of water, weak acids, and weak bases****Page: 57 Difficulty: 1 Ans: D**

The aqueous solution with the lowest pH is:

- A) 0.01 M HCl.
- B) 0.1 M acetic acid ( $pK_a = 4.86$ ).
- C) 0.1 M formic acid ( $pK_a = 3.75$ ).
- D) 0.1 M HCl.
- E)  $10^{-12}$  M NaOH.

**12. Ionization of water, weak acids, and weak bases****Page: 57 Difficulty: 1 Ans: D**

The aqueous solution with the highest pH is:

- A) 1 M HCl.
- B) 1 M  $NH_3$  ( $pK_a = 9.25$ ).
- C) 0.5 M  $NaHCO_3$  ( $pK_a = 3.77$ ).
- D) 0.1 M NaOH.
- E) 0.001 M NaOH.

**13. Ionization of water, weak acids, and weak bases****Page: 57 Difficulty: 2 Ans: B**Phosphoric acid is tribasic, with  $pK_a$ 's of 2.14, 6.86, and 12.4. The ionic form that predominates at pH 3.2 is:

- A)  $H_3PO_4$ .
- B)  $H_2PO_4^-$ .
- C)  $HPO_4^{2-}$ .
- D)  $PO_4^{3-}$ .
- E) none of the above.

**14. Buffering against pH changes in biological systems****Page: 59-60 Difficulty: 2 Ans: E**

Which of the following statements about buffers is true?

- A) A buffer composed of a weak acid of  $pK_a = 5$  is stronger at pH 4 than at pH 6.
- B) At pH values lower than the  $pK_a$ , the salt concentration is higher than that of the acid.
- C) The pH of a buffered solution remains constant no matter how much acid or base is added to the solution.
- D) The strongest buffers are those composed of strong acids and strong bases.
- E) When  $pH = pK_a$ , the weak acid and salt concentrations in a buffer are equal.

**15. Buffering against pH changes in biological systems****Page: 59-61      Difficulty: 3      Ans: D**

A compound has a  $pK_a$  of 7.4. To 100 mL of a 1.0 M solution of this compound at pH 8.0 is added 30 mL of 1.0 M hydrochloric acid. The resulting solution is pH:

- A) 6.5
- B) 6.8
- C) 7.2
- D) 7.4
- E) 7.5

**16. Buffering against pH changes in biological systems****Page: 60-61      Difficulty: 2      Ans: E**

The Henderson-Hasselbalch equation:

- A) allows the graphic determination of the molecular weight of a weak acid from its pH alone.
- B) does not explain the behavior of di- or tri-basic weak acids
- C) employs the same value for  $pK_a$  for all weak acids.
- D) is equally useful with solutions of acetic acid and of hydrochloric acid.
- E) relates the pH of a solution to the  $pK_a$  and the concentrations of acid and conjugate base.

**17. Buffering against pH changes in biological systems****Page: 60-61      Difficulty: 2      Ans: D**

Consider an acetate buffer, initially at the same pH as its  $pK_a$  (4.76). When sodium hydroxide (NaOH) is mixed with this buffer, the:

- A) pH remains constant.
- B) pH rises more than if an equal amount of NaOH is added to an acetate buffer initially at pH 6.76.
- C) pH rises more than if an equal amount of NaOH is added to unbuffered water at pH 4.76.
- D) ratio of acetic acid to sodium acetate in the buffer falls.
- E) sodium acetate formed precipitates because it is less soluble than acetic acid.

**18. Buffering against pH changes in biological systems****Page: 60-61      Difficulty: 3      Ans: C**

A compound is known to have a free amino group with a  $pK_a$  of 8.8, and one other ionizable group with a  $pK_a$  between 5 and 7. To 100 mL of a 0.2 M solution of this compound at pH 8.2 was added 40 mL of a solution of 0.2 M hydrochloric acid. The pH changed to 6.2. The  $pK_a$  of the second ionizable group is:

- A) The pH cannot be determined from this information.
- B) 5.4
- C) 5.6
- D) 6.0
- E) 6.2

**19. Buffering against pH changes in biological systems****Page: 60–61      Difficulty: 2      Ans: C**

Three buffers are made by combining a 1 M solution of acetic acid with a 1 M solution of sodium acetate in the ratios shown below.

	1 M acetic acid	1 M sodium acetate
Buffer 1:	10 mL	90 mL
Buffer 2:	50 mL	50 mL
Buffer 3:	90 mL	10 mL

Which of these statements is true of the resulting buffers?

- A) pH of buffer 1 < pH of buffer 2 < pH of buffer 3
- B) pH of buffer 1 = pH of buffer 2 = pH of buffer 3
- C) pH of buffer 1 > pH of buffer 2 > pH of buffer 3
- D) The problem cannot be solved without knowing the value of  $pK_a$ .
- E) None of the above.

**20. Buffering against pH changes in biological systems****Page: 61–63      Difficulty: 3      Ans: A**

A 1.0 M solution of a compound with 2 ionizable groups ( $pK_a$ 's = 6.2 and 9.5; 100 mL total) has a pH of 6.8. If a biochemist adds 60 mL of 1.0 M HCl to this solution, the solution will change to pH:

- A) 5.60
- B) 8.90
- C) 9.13
- D) 9.32
- E) The pH cannot be determined from this information.

**21. Water as a reactant****Page: 65      Difficulty: 3      Ans: E**

In which reaction below does water *not* participate as a reactant (rather than as a product)?

- A) Conversion of an acid anhydride to two acids.
- B) Conversion of an ester to an acid and an alcohol.
- C) Conversion of ATP to ADP.
- D) Photosynthesis
- E) Production of gaseous carbon dioxide from bicarbonate.

**22. The fitness of the aqueous environment for living organisms****Pages: 65–66      Difficulty: 1      Ans: E**

Which of the following properties of water does *not* contribute to the fitness of the aqueous environment for living organisms?

- A) Cohesion of liquid water due to hydrogen bonding.
- B) High heat of vaporization.
- C) High specific heat.
- D) The density of water is greater than the density of ice.
- E) The very low molecular weight of water.

## Short Answer Questions

### 23. Weak interactions in aqueous systems

**Page:** 43–51      **Difficulty:** 2

Name and briefly define four types of noncovalent interactions that occur between biological molecules.

**Ans:** (1) Hydrogen bonds: weak electrostatic attractions between one electronegative atom (such as oxygen or nitrogen) and a hydrogen atom covalently linked to a second electronegative atom; (2) electrostatic interactions: relatively weak charge-charge interactions (attractions of opposite charges, repulsions of like charges) between two ionized groups; (3) hydrophobic interactions: the forces that tend to bring two hydrophobic groups together, reducing the total area of the two groups that is exposed to surrounding molecules of the polar solvent (water); (4) van der Waals interactions: weak interactions between the electric dipoles that two close-spaced atoms induce in each other.

### 24. Weak interactions in aqueous systems

**Page:** 46–48      **Difficulty:** 3

Explain the fact that ethanol ( $\text{CH}_3\text{CH}_2\text{OH}$ ) is more soluble in water than is ethane ( $\text{CH}_3\text{CH}_3$ ).

**Ans:** Ethanol can form hydrogen bonds with water molecules, but ethane cannot. When ethanol dissolves, the decrease in the system's entropy that results from formation of ordered arrays of water around the  $\text{CH}_3\text{CH}_2^-$  group is partly compensated by the favorable interactions (hydrogen bonds) of the hydroxyl group of ethanol with water molecules. Ethane cannot form such hydrogen bonds.

### 25. Weak interactions in aqueous systems

**Page:** 46–48      **Difficulty:** 3

Explain the fact that triethylammonium chloride ( $(\text{CH}_3\text{CH}_2)_3\text{N}\cdot\text{HCl}$ ) is more soluble in water than is triethylamine ( $(\text{CH}_3\text{CH}_2)_3\text{N}$ ).

**Ans:** The positive charge on the N atom in triethylammonium chloride is more polar than the uncharged N atom in triethylamine. This increased polarity leads to stronger interactions with water, leading to increased solubility.

### 26. Weak interactions in aqueous systems

**Page:** 48      **Difficulty:** 3

Explain with an appropriate diagram why amphipathic molecules tend to form micelles in water. What force drives micelle formation?

**Ans:** Micelle formation minimizes the area of the hydrophobic part of amphipathic molecules that contacts the polar solvent, water. Hydrophobic interactions between hydrophobic moieties are the driving force for micelle formation. When amphipathic molecules form micelles in water, the entropy decrease due to the formation of ordered arrays of water molecules around the hydrophobic moieties is minimized. (See Fig. 2-7, p. 48.)

**27. Weak interactions in aqueous systems****Pages: 51–52      Difficulty: 2**

(a) Briefly define “isotonic,” “hypotonic,” and “hypertonic” solutions. (b) Describe what happens when a cell is placed in each of these types of solutions.

**Ans:** (a) An isotonic solution has the same osmolarity as the solution to which it is being compared. A hypotonic solution has a lower osmolarity than the solution to which it is being compared. A hypertonic solution has a higher osmolarity than the solution to which it is being compared. (b) Higher osmolarity results in osmotic pressure, which generally leads to movement of water across a membrane. In an isotonic solution, in which the osmolarity of the solution is the same as the cell cytoplasm, there will be no net water movement. In a hypotonic solution, water will move into the cell, causing the cell to swell and possibly burst. In a hypertonic solution, water will move out of the cell and it will shrink.

**28. Ionization of water, weak acids, and weak bases****Page: 57 Difficulty: 1**

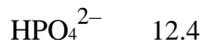
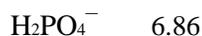
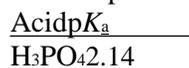
For each of the pairs below, circle the conjugate base.



**Ans:**  $\text{RCOO}^-$ ,  $\text{RNH}_2$ ,  $\text{H}_2\text{PO}_4^-$ ,  $\text{HCO}_3^-$

**29. Ionization of water, weak acids, and weak bases****Page: 57 Difficulty: 2**

Phosphoric acid ( $\text{H}_3\text{PO}_4$ ) has three dissociable protons, with the  $\text{p}K_a$ 's shown below. Which form of phosphoric acid predominates in a solution at pH 4? Explain your answer.



**Ans:** At pH 4, the first dissociable proton ( $\text{p}K_a = 2.14$ ) has been titrated completely, and the second ( $\text{p}K_a = 6.86$ ) has just started to be titrated. The dominant form at pH 4 is therefore  $\text{H}_2\text{PO}_4^-$ , the form with one dissociated proton (see Fig. 2-15).

**30. Ionization of water, weak acids, and weak bases****Page: 58–59      Difficulty: 1**

Define  $\text{p}K_a$  for a weak acid in the following two ways: (1) in relation to its acid dissociation constant,  $K_a$ , and (2) by reference to a titration curve for the weak acid.

**Ans:** (1)  $\text{p}K_a = -\log K_a$ ; (2) See Fig. 2-17, p. 59;  $\text{p}K_a$  is the value of pH at the inflection point in a plot of pH vs. extent of titration of the weak acid. At the  $\text{p}K_a$ , the concentration of ionized acid equals the concentration of un-ionized acid.

**31. Buffering against pH changes in biological systems****Page: 58–60      Difficulty: 2**

Give the general Henderson-Hasselbalch equation and sketch the plot it describes (pH against amount of NaOH added to a weak acid). On your curve label the  $pK_a$  for the weak acid, and indicate the region in which the buffering capacity of the system is greatest.

**Ans:** The inflection point, which occurs when the weak acid has been exactly one half titrated with NaOH, occurs at a pH equal to the  $pK_a$  of the weak acid. The region of greatest buffering capacity (where the titration curve is flattest) occurs at pH values of  $pK_a \pm 1$ . (See Fig. 2-17, p. 59.)

**32. Buffering against pH changes in biological systems****Pages: 59–60      Difficulty: 3**

Draw the titration curve for a weak acid, HA, whose  $pK_a$  is 3.2. Label the axes properly. Indicate with an arrow where on the curve the ratio of salt ( $A^-$ ) to acid (HA) is 3:1. What is the pH at this point?

**Ans:** The plot of pH vs. added base should have the general shape of those shown in Fig. 2-17, p. 59, with the midpoint of the titration (inflection point) at pH 3.2. The ratio of  $A^-$  to HA is 3 when 0.75 equivalents of base have been added. From the Henderson-Hasselbalch equation, the pH at this point can be calculated:

$$pH = pK_a + \log \frac{[\text{conjugate base}]}{[\text{acid}]} = 3.2 + \log 3 = 3.2 + 0.48 = 3.68$$

**33. Buffering against pH changes in biological systems****Page: 61–62      Difficulty: 2**

What is the pH of a solution containing 0.2 M acetic acid ( $pK_a = 4.7$ ) and 0.1 M sodium acetate?

**Ans:**

$$\begin{aligned} pH &= pK_a + \log \frac{[\text{conjugate base}]}{[\text{acid}]} = 4.7 + \log (0.1/0.2) \\ &= 4.7 - 0.3 = 4.4 \end{aligned}$$

**34. Buffering against pH changes in biological systems****Page: 61–62      Difficulty: 2**

You have just made a solution by combining 50 mL of a 0.1 M sodium acetate solution with 150 mL of 1 M acetic acid ( $pK_a = 4.7$ ). What is the pH of the resulting solution?

**Ans:**

$$\begin{aligned} pH &= pK_a + \log \frac{[\text{conjugate base}]}{[\text{acid}]} = 4.7 + \log (5/150) \\ &= 4.7 - 1.48 = 3.22 \end{aligned}$$

**35. Buffering against pH changes in biological systems****Page: 61–62 Difficulty: 2**For a weak acid with a  $pK_a$  of 6.0, show how you would calculate the ratio of acid to salt at pH 5.**Ans:**

$$\begin{aligned} \text{pH} &= pK_a + \log \frac{[\text{conjugate base}]}{[\text{acid}]}, \text{ so } pK_a - \text{pH} = -\log \frac{[\text{conjugate base}]}{[\text{acid}]} \\ &= \log \frac{[\text{acid}]}{[\text{conjugate base}]} \\ 6.0 - 5.0 &= \log \frac{[\text{acid}]}{[\text{conjugate base}]}; \frac{[\text{acid}]}{[\text{conjugate base}]} = \text{antilog } 1 = 10 \end{aligned}$$

**36. Buffering against pH changes in biological systems****Page: 61-62 Difficulty: 3**Suppose you have just added 100 mL of a solution containing 0.5 mol of acetic acid per liter to 400 mL of 0.5 M NaOH. What is the final pH? (The  $pK_a$  of acetic acid is 4.7.)

**Ans:** Addition of 200 mmol of NaOH (400 mL 0.5 M) to 50 mmol of acetic acid (100 mL 0.5 mM) completely titrates the acid so that it can no longer act as a buffer and leaves 150 mmol of NaOH dissolved in 500 mL, an  $[\text{OH}^-]$  of 0.3 M. Given  $[\text{OH}^-]$ ,  $[\text{H}^+]$  can be calculated from the water constant:

$$[\text{H}^+][\text{OH}^-] = 10^{-14}$$

$$[\text{H}^+] = 10^{-14} \text{ M}^2 / 0.3 \text{ M}$$

$$\begin{aligned} \text{pH is, by definition, } &\log (1/[\text{H}^+]) \\ \text{pH} &= \log (0.3 \text{ M} / 10^{-14} \text{ M}^2) = 12.48. \end{aligned}$$

**37. Buffering against pH changes in biological systems****Page: 61-62 Difficulty: 2**A weak acid HA, has a  $pK_a$  of 5.0. If 1.0 mol of this acid and 0.1 mol of NaOH were dissolved in one liter of water, what would the final pH be?

**Ans:** Combining 1 mol of weak acid with 0.1 mol of NaOH yields 0.9 mol of weak acid and 0.1 mol of salt.

$$\text{pH} = pK_a + \log \frac{[\text{conjugate base}]}{4.05 [\text{acid}]} = 5.0 + \log (0.1/0.9) =$$

**38. Water as a reactant****Page: 65 Difficulty: 1**

Give an example of a biological reaction in which water participates as a reactant and a reaction in which it participates as a product.

**Ans:** See p. 65, for examples such as condensation and hydrolysis reactions.

**39. The fitness of the aqueous environment for living organisms****Pages: 65–66 Difficulty: 1**

If ice were denser than water, how would that affect life on earth?

**Ans:** Ice formed at the surface of bodies of water would sink, hence streams, ponds, lakes, and so on would freeze from the bottom up. With a reservoir of ice at the bottom they would be perpetually cold, and in the limit they would freeze solid, precluding life as we know it.