

# **Biochemistry**

Water Chemistry, Titration, and Macromolecules

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# 1 Introduction

## 1.1 Opening Statement

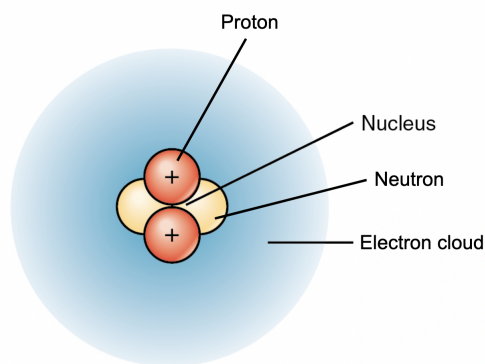
Biochemistry is a captivating blend of biology and chemistry, and certain principles found in biochemistry are needed to fully understand the background of life's most complex biological processes. Water's unique chemical properties, amino acid interactions with solutions, basic biological macromolecules, and more essential concepts all play their own part in understanding biology and the phenomenon of life.

# 2 Review of Chemistry

In order to understand the following contents of this handout, a basic understanding of chemistry is required. This section will provide a brief explanation of its most important concepts. Electrons will be a major focus of this review.

## 2.1 Structure of an Atom

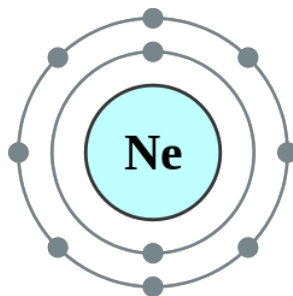
All of matter is made up of elements, each with their own properties. These elements exist as atoms, tiny particles consisting of **protons**, **neutrons**, and **electrons**. An atom's **nucleus**, or center, is made of protons and neutrons. On the other hand, electrons reside in the **electron cloud**, the area surrounding an atom's nucleus. Protons have a **positive electric charge**, while electrons have a **negative electric charge**. Neutrons, as inferred from its name, are neutral.



**Figure 2.1** A diagram of an atom and its constituent parts. (Source: OpenStax College)

### 2.1.1 Electron Shells

Electrons have varying levels of energy, causing them to reside in different **electron shells** inside of the electron cloud. Each electron shell can hold a certain number of electrons, and this greatly affects an atom's reactivity and properties. An atom is most stable when its **valence electron shell**, or outermost electron shell, is holding electrons to its capacity. For example, the element neon has a filled valence electron shell and is thus very inert and non-reactive.



**Figure 2.2** A model of an atom of the element neon, showing electrons (dots) and electron shells (circles). (Source: Greg Robson)

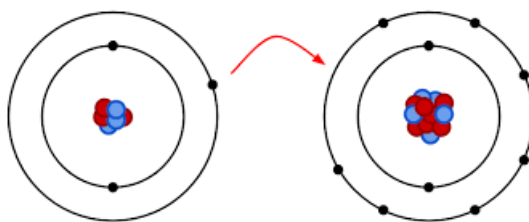
## 2.2 Electronegativity and Bonding

In contrast to neon's naturally filled valence electron shell, an atom of an element such as fluorine has one electron missing from its valence electron shell. This brings us to the concept of **electronegativity**, or the measure of an atom's ability to pull electrons to itself. Fluorine is not stable and is very reactive due to how close it is to reaching a filled valence electron shell, and it will rip electrons away from other atoms in order to fill said shell. Fluorine thus has a very high electronegativity value.

Elements such as lithium are also very reactive, but are not very electronegative. Instead, their valence electron shells are filled with very few electrons. Instead of pulling other electrons to fill it, these elements want to lose electrons to reveal a lower, filled valence electron shell.

### 2.2.1 Ionic Bonding

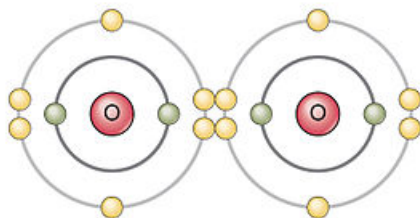
When atoms from two elements such as fluorine and lithium interact, the lithium wants to lose its electrons to reach stability, and the fluorine wants to gain electrons to reach stability. When reacted, exactly this happens: lithium gives up its electron to fluorine, leading to both elements having a complete valence electron shell. This bonds the two together through **ionic bonding** to create the compound lithium fluoride.



**Figure 2.3** An atom of the element lithium (left) shown giving an electron to an atom of the element fluorine (right), leading to ionic bonding and completion of both atoms' valence electron shells.

### 2.2.2 Covalent Bonding

In lithium and fluorine, the differences in electronegativity are great enough to warrant an atom of fluorine entirely ripping away an electron. However, when two atoms with similar electronegativity values react, something different happens. Take two atoms of oxygen for example, both having the exact same electronegativity. Both atoms need two electrons to form a complete valence electron shell, however pulling electrons away from the other would not work, as it would lead to only one oxygen atom with a complete shell. Instead, the oxygen atoms share each others' electrons, binding them together through **covalent bonding**.



**Figure 2.4** Two atoms of oxygen exhibiting covalent bonding. Through sharing two electrons with the other atom, both atoms have a completed valence electron shell. (Source: OpenStax College)

## 3 Chemistry of Water

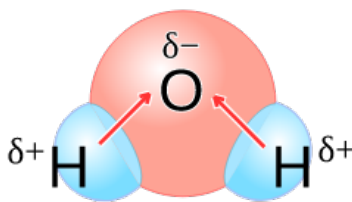
Water has many unique properties, many of which make life possible. In this section, we'll be exploring these properties and the chemistry behind why they occur.

### 3.1 Hydrogen Bonding

In a water molecule ( $\text{H}_2\text{O}$ ), there are two hydrogen atoms ( $\text{H}_2$ ) and one oxygen atom ( $\text{O}$ ). Both hydrogen and oxygen have high electronegativity values, but oxygen has a slightly greater electronegativity.

This difference in electronegativity does not warrant ionic bonding; the atoms still share electrons in covalent bonding. However, it does lead to an unequal sharing of electrons. When an oxygen atom is covalently bonded to a hydrogen atom, the shared electrons are pulled more towards the oxygen atom due its greater electronegativity.

Remember that electrons have a negative charge associated with them. Since the electrons are pulled closer to the oxygen atom, the oxygen end of the compound is slightly more negative than the hydrogen ends. This makes it a **polar molecule**, a molecule with both positive and negative charges. Water has a partial negative charge  $\delta^-$  near the oxygen end, and a partial positive charge  $\delta^+$  near the hydrogen ends. Although this slight difference in charges doesn't seem too important when looking at one isolated molecule of water, its more pronounced effects show when we look at a group of water molecules together.



**Figure 3.1** Partial charges in a water molecule (red arrows represent pulling of electrons).

With more than one water molecule, **intermolecular forces** come into play. Opposite electric charges attract, and one water molecule's  $\delta^-$  charge will create a weak bond with another water molecule's  $\delta^+$  charge. The intermolecular forces between the hydrogen atom end of one molecule and the oxygen atom end of another are called **hydrogen bonds** (not to be confused with ionic or covalent bonds, as those are **intramolecular forces**).

## 3.2 Unique Properties of Water

### 3.2.1 Cohesion and Adhesion

Due to this hydrogen bonding, water molecules tend to stick together, leading to **cohesion**. Water can also stick to other substances with the same hydrogen bonding, leading to **adhesion**. Both adhesion and cohesion come together to make some of life's processes possible; for example, in trees, adhesion between water molecules and plant cell walls resists the downward pull of gravity. Cohesion between multiple water molecules maintains a stable column of water. These two properties allow water to move up the comparatively massive tree trunk and reach leaves at even the tallest of canopies.

### 3.2.2 Surface Tension

**Surface tension** is another property of water caused by hydrogen bonding. Since water molecules are hydrogen bonded to each other, but not the surrounding atmosphere or air particles, it is relatively difficult to break through the surface of a body of water; all the surface water molecules adhere strongly to only each other and the water molecules below them. This property is what allows bugs such as water striders to walk across the surface of a pond without sinking in; the surface tension is unbroken and holds the bug up.

### 3.2.3 Specific Heat and Temperature

**Kinetic energy** is the energy of motion. **Thermal energy** is a form of kinetic energy, and thus anything with a high **temperature**, has a high amount of kinetic energy. In fact, the scientific definition of temperature is the average kinetic energy of a body of matter.

Remember that hydrogen bonds between water molecules keep water molecules close to each other. When thermal energy in the form of kinetic energy is introduced, the water molecules begin to move and increase in temperature. However, in order for this to happen, the hydrogen bonds between water molecules must be broken—enough thermal energy must be added to first push apart the water molecules enough to break the hydrogen bonds before they can be fully separated.

Due to the strength of hydrogen bonds, water requires a large amount of thermal energy in order to increase its temperature. Thus, water has a high **specific heat**, a measure of how much thermal energy must be absorbed to change a substance's temperature. Many other commonly occurring liquid compounds cannot match water's high specific heat (due to hydrogen bonds being the strongest type of intermolecular force), and we can see why this is so important when bringing up real-world examples.

Large bodies of water can act as temperature buffers—for example, the Earth's oceans can absorb a massive amount of thermal energy from the sun without significantly increasing its temperature, and also maintain that heat without the sun, leading to marine ecosystems being able to flourish in both day and night.

### 3.2.4 Evaporative Cooling

Water, for the same reason it has a high specific heat, requires a large amount of thermal energy in order to evaporate. This is again, due to the hydrogen bonds between water molecules needing to break before water molecules can separate and evaporate.

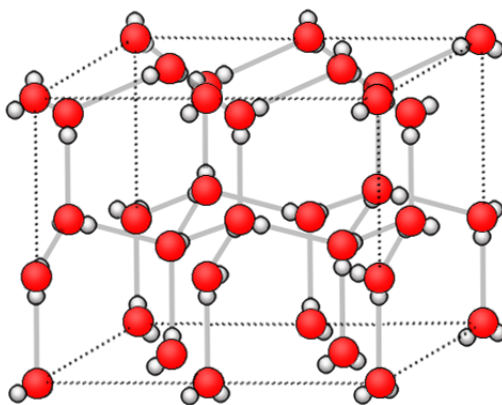
When thermal energy is introduced, a water molecule eventually evaporates, taking along with it the relatively large amount of thermal energy as it moves away from the initial body of water. This leads to **evaporative cooling**, where vaporization of water leads to overall cooling of a greater body of water. This is important as it also helps with water's role as a temperature buffer, being able to cool down lakes and oceans even while the sun is beating down on the water's surface.

### 3.2.5 Ice Density

Water has a very unique property in which its solid form, ice, is less dense than its liquid form. This is, once again, due to hydrogen bonding. Hydrogen bonds, although the strongest of all the intermolecular forces, are not very strong when compared to other natural forces—they constantly break and reform while water is in a liquid state of matter.

At any moment, there are many chains of water molecules hydrogen bonded together, however not to their neighbors, and this allows water molecules to slip between each other.

However, when the temperature decreases and water molecules lose thermal energy and begin to slow down, more and more hydrogen bonds are able to form. Comparatively, the water molecule chains become larger and larger, until the water molecules are held in a crystal lattice. In this case, there is no slippage of water molecules between each other; every water molecule is positioned equidistantly from the next.



**Figure 3.2** Water molecules held by hydrogen bonds (solid grey lines) in a crystal lattice. Note how each water molecule is equidistant from the others, and there is no slippage of water molecules unlike what one would see when observing liquid water. (Source: IgniX)

Imagine throwing a bunch of objects into a box, they will fill up empty spaces and slide into nooks and crevices. Then, imagine carefully positioning each object so that there is empty space between them. This is analogous to water molecules as a liquid and as a solid.

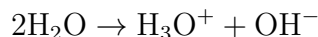
This phenomenon is important for many reasons. For example, many polar environments simply would not exist if ice did not float on water: there would be no icebergs or glaciers.

## 4 Titration

Here, the basics of acidity and its important roles in allowing life will be covered.

### 4.1 Acids and Bases

When two water molecules are hydrogen bonded together, sometimes a hydrogen atom from one molecule leaves and bonds with the other molecule, changing two water molecules into one hydronium and hydroxide ion:



In this process, the hydrogen atom does not bring any electrons with it, only its singular proton. This is the reason that the hydronium ion has a positive charge (due to gaining a positively-charged hydrogen atom) and the hydroxide ion has a negative charge (due to losing a positively-charged hydrogen atom). The hydronium ion is commonly shortened to simply  $\text{H}^+$ , however know that it represents  $\text{H}_3\text{O}^+$ .

The relationship between  $\text{H}^+$  and  $\text{OH}^-$  in a solution is what acidity is based off of. In a pure solution of water, there is a roughly equal amount of  $\text{H}^+$  and  $\text{OH}^-$  at any given time. It is also important to note that this is a **dynamic equilibrium**, meaning there is constant fluctuation of  $\text{H}^+$  and  $\text{OH}^-$ , however when one increases in concentration more than the other, the reactions tilt to the other side to reach back to equilibrium.

- An **acid** is a molecule that when added to water, dissolves and adds  $\text{H}^+$  to the solution, resulting in a greater concentration of  $\text{H}^+$  compared to  $\text{OH}^-$ .



- A **base** is a molecule that when added to water, dissolves and adds  $\text{OH}^-$  to the solution, resulting in a greater concentration of  $\text{OH}^-$  compared to  $\text{H}^+$ .

In order to more easily quantify acidity, the **pH scale** is used. The pH of a solution is determined by the equation:

$$\text{pH} = -\log[\text{H}^+]$$

where  $[\text{H}^+]$  is the concentration of  $\text{H}^+$  in **moles** ( $6.022^{23}$  molecules) per litre. In a neutral solution of water, there are  $10^{-7}$  moles of  $\text{H}^+$  per litre, so we use this equation to determine:

$$\text{pH} = -\log[10^{-7}] = 7$$

If an acid is added, the concentration of  $\text{H}^+$  increases, and the pH decreases. The opposite is true for bases, making any solution with a pH below 7 **acidic**, any solution with a pH above 7 **basic**, and any solution with a pH of 7 neutral.

Certain molecules can act in acid-base pairs, for example the two molecules acetic acid and acetate, which can give and take  $\text{H}^+$  ions respectively (taking  $\text{H}^+$  ions is essentially the same as adding  $\text{OH}^-$  ions). When an acid donates  $\text{H}^+$  ions, it becomes its **conjugate base**, and is able to now up a  $\text{H}^+$  ion. a Acetic acid can give  $\text{H}^+$  ions, turning into acetate, and acetate can take  $\text{H}^+$  ions, turning back into acetic acid. Acetate is thus known as the conjugate base of acetic acid. This property to be able to convert between both an acid and a base allows **buffer solutions** to be made, or solutions which resist changes in pH.

In an acidic solution, there will be a greater concentration of acetic acid than acetate. In a basic solution, there will be a greater concentration of acetate than acetic acid. This is due to something called Le Chatelier's principle, stated as:

- **Le Chatelier's principle:** When a system at equilibrium is subjected to a change in temperature, pressure, concentration of reactants or products, or volume, the system will adjust itself in such a way as to counteract the effect of the change and restore equilibrium.

When acetic acid is added to a basic solution at equilibrium, it will dissociate due to there already being a greater concentration of  $\text{OH}^-$  compared to  $\text{H}^+$ . Dissociation will increase the concentration of  $\text{H}^+$ , bringing it closer to equilibrium. Dissociation will also lead to the acetic acid becoming acetate. The opposite is true for acetate being added to an acidic solution.

How about when acetic acid is added to an acidic solution? There will be significantly less dissociation due to there already being an imbalance,  $\text{H}^+$  is in greater concentration than  $\text{OH}^-$ . Dissociation would not lead to greater equilibrium, and thus its effects are not as pronounced.

#### 4.1.1 $K_a$ and $\text{p}K_a$

Some acids, such as hydrochloric acid ( $\text{HCl}$ ), completely dissolve and dissociate with its hydrogen atoms when added to a solution:



These acids are known as **strong acids** (the same concept is true for **strong bases**). Other acids do not completely dissociate with their hydrogen atoms, known as **weak acids** (again, same concept for **weak bases**). Acetic acid and acetate are examples of a weak acid and a weak base.

All acids have a  $K_a$  value, or an **acid dissociation constant**, measuring how weak or strong an acid is. An acid is stronger the more it is able to release  $H^+$  ions, thereby increasing the acidity of a solution. The greater the  $K_a$ , the stronger the acid, the more it is able to give off  $H^+$  ions, and the more it can decrease pH.  $K_a$  is calculated using the equation:

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

where  $[A^-]$  is the conjugate base concentration,  $[H^+]$  is the  $H^+$  concentration, and  $[HA]$  is the conjugate acid concentration.

$pK_a$  is  $K_a$  on a logarithmic scale.  $pK_a$  is useful as it allows the relation of  $K_a$  to pH, and can be calculated using the equation:

$$pK_a = -\log[K_a]$$

The **Henderson-Hasselbalch equation** builds upon  $pK_a$  and allows the mathematical relation of pH to  $pK_a$ , where:

$$pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right)$$

A buffer solution is most effective when both concentrations of a weak acid and its conjugate base are equal. When said concentrations are equal,  $[A^-]/[HA] = 1$ . Using the Henderson-Hasselbalch equation results in the equation  $pH = pK_a + \log(1)$ , and simplifying further results in

$$pH = pK_a$$

This tells us that a buffer solution is most effective when  $pH = pK_a$ .

**Example 4.1** (USABO Open Exam 2010) The Henderson-Hasselbalch equation describes the relationship between pH,  $pK_a$  and the ratio of base to acid concentration:

$$pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right)$$

It is useful when determining the amount of weak acid and weak base to weigh out for a buffer system. It is also useful to determine percent of a functional group in its ionized state. If a carboxylic acid has a  $pK_a$  of 3.9, what percentage of the functional group will carry a negative charge at pH 4.9?

- (A) 10 percent
- (B) 90.9 percent
- (C) 9.09 percent
- (D) 50 percent
- (E) 18.18 percent

**Solution:** In this question, we are being asked to find the percentage of the functional group with a negative charge at a specific pH. Recall that conjugate bases are negatively charged—they have given off their  $\text{H}^+$  ion, losing a positively charged ion, leading to it being negatively charged. Thus we are finding the percentage of conjugate base in the solution, or the variable  $[\text{A}^-]$ . Isolating  $[\text{A}^-]$  using the Henderson-Hasselbalch equation results in the new equation

$$[\text{A}^-] = ([\text{HA}])(10)^{\text{pH}-\text{p}K_a}$$

Plugging in the provided values of 4.9 and 3.9 and for pH and  $\text{p}K_a$  results in the equation  $[\text{A}^-] = ([\text{HA}])(10)$ . We know that the percentages of weak acid and conjugate base add up to 100 percent, so we can substitute  $[\text{HA}]$  with  $1 - [\text{A}^-]$ , leading to the new equation

$$[\text{A}^-] = 10 - 10([\text{A}^-])$$

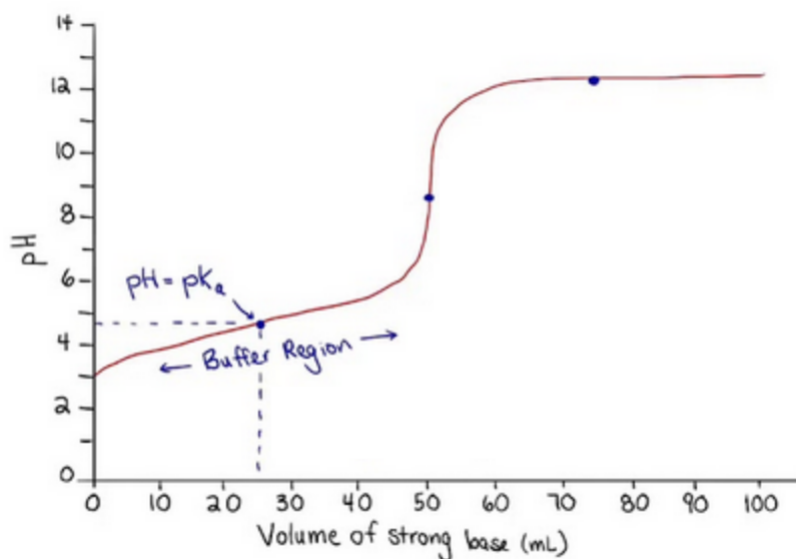
Simplifying for  $[\text{A}^-]$  leads to a value of 0.909, or 90.9 percent. The answer is thus B.

#### 4.1.2 Titration Curves

$\text{p}K_a$  can be effectively visualized and determined through an experimental technique known as **titration**, where the pH of a buffer solution is measured as a strong base is constantly added to increase the pH. Recall that strong bases, similar to strong acids, completely dissociate when dissolved in a solution of water. However, instead of dissociating and releasing a  $\text{H}^+$  ion, strong bases dissociate and release  $\text{OH}^-$  ions. A common strong base used in titration is sodium hydroxide ( $\text{NaOH}$ ).

Titration can be visualized through **titration curves**. Titration curves are graphs visualizing quantitative data from a titration, plotting **pH of a solution** and **volume of a strong base added to a solution**.

On a titration curve, the point where  $\text{pH} = \text{p}K_a$  is known as the **half-equivalence point**, due to the buffer solution containing half conjugate bases and half weak acids. Recall that buffer solutions are most effective around the pH where  $\text{pH} = \text{p}K_a$ . On a titration curve, this area of effectiveness is known as a **buffer region**, and is signified by a flat region in the graph (as strong base is added, there is not much increase in pH due to the effects of the buffer, so the slope of the graph in a buffer region is low).



**Figure 4.1** A titration curve of a buffer solution. Note the buffer region, and the half-equivalence point. (Source: Kognity HL Biology)

## 4.2 Amino Acid Titration

**Amino acids** are biological molecules containing **amine** and **carboxylic acid**.

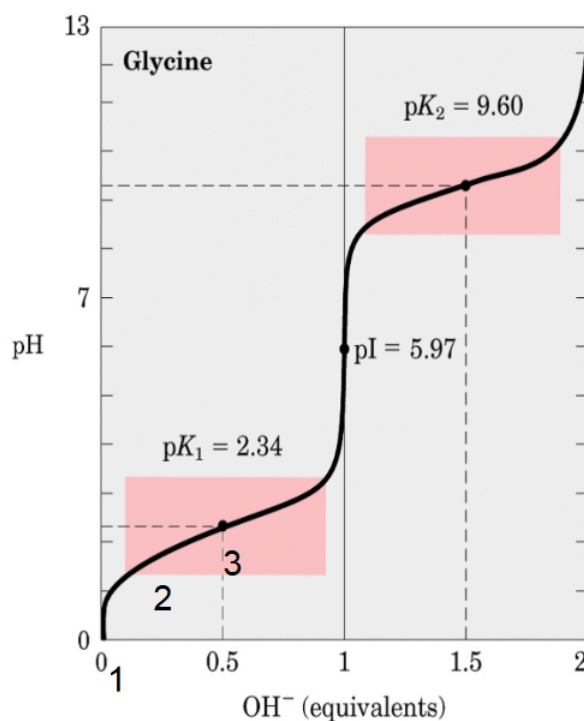
- Recall that carboxylic acid is able to donate  $H^+$  ions to a solution, acting as an acid and having a conjugate base form.
- The amine in an amino acid is able to take up  $H^+$  ions from a solution, acting as a base and having a conjugate acid form.

Amino acids can have multiple configurations due to containing two groups which can both act as buffers. For example, in very low pH solutions, both the carboxylic acid group and the amine group in the amino acid act as bases. As the pH of the solution increases, some of the carboxylic acid groups may begin to dissociate, releasing  $H^+$  ions and acting as an acid. As the pH increases further, some of the amine groups may begin to dissociate as well.

Recall that buffer solutions work using concentrations; there is no pH in which, for example, all carboxylic acid groups suddenly dissociate—it is a gradual scale, for example the concentration of amino acids with dissociated carboxylic acid groups compared to amino acids without dissociated carboxylic acid groups increases as pH increases.

Amino acids are **diprotic**, meaning they can give and take up to two  $H^+$  ions. Acetic acid can give and take up to one  $H^+$  ion, and is thus **monoprotic**.

Due to having two groups which can both give and take  $H^+$  ions, amino acids have two  $pK_a$  values, one for the carboxylic acid group and one for the amine group. Having two  $pK_a$  values means that there are two points where  $pH = pK_a$ , meaning that diprotic amino acids can function as effective buffers in two ranges on the pH scale. The two  $pK_a$  values can be notated as  $pK_{a_1}$  and  $pK_{a_2}$ .



**Figure 4.2** A titration curve of glycine, a diprotic amino acid. Note the two  $pK_a$  values, and the resulting two effective buffer regions (highlighted red). (Source: Kansas State University)

The **isoelectric point (pI)** of an amino acid is the pH at which an amino acid does not give or take any  $H^+$  ions, and is electrically neutral. The amino acid has both a positive and a negative charge due to this, and it is thus known as a **zwitterion**. At the pI of an amino acid, all molecules of that amino acid are in the zwitterion form. The pI of an amino acid can be calculated by finding the average of both  $pK_a$  values.

It is important to understand how an amino acid interacts with solutions with varying pH levels as it allows the prediction of an amino acid's structure and how it'll change the environment it's in. Amino acids are what make up proteins in cells, and by understanding how an amino acid will react in a solution allows prediction of how a resulting protein will fold and behave.

**Example 4.1** (Adapted from CMU Biochemistry I Exam 2005) Diprotic amino acids such as glycine, can act as effective acid-base buffers:

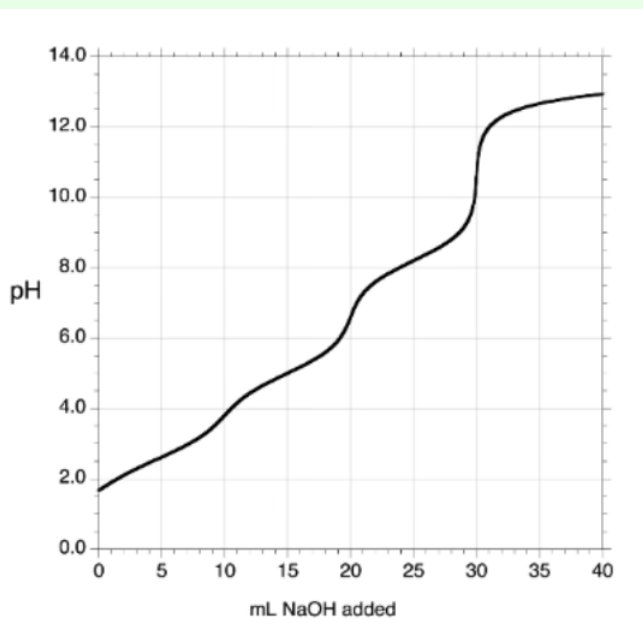
- (A) Only in combination with other diprotic bases
- (B) At pH values around any of their  $pK_a$  values
- (C) At pH values around neutrality
- (D) At pH values around their singular  $pK_a$  value
- (E) At the pH value of the average of their  $pK_a$  values

**Solution:** Recall that amino acids are diprotic and have two groups which can each give and take  $H^+$  ions, each with their own disassociation constant  $K_a$ , and thus two  $pK_a$  values. Recall that molecules are most effective at being acid-base buffers when the concentration of their weak acid and conjugate base forms are equal. When plugging equal concentrations of acid and conjugate base forms into the Henderson-Hasselbalch equation, we get the equation  $pH = pK_a$ , meaning that molecules are most effective as buffers when their  $pK_a$  values are equal or around the pH of the solution. Combining all of this information together leads us to answer B.

#### 4.2.1 Triprotic and Polyprotic Amino Acids

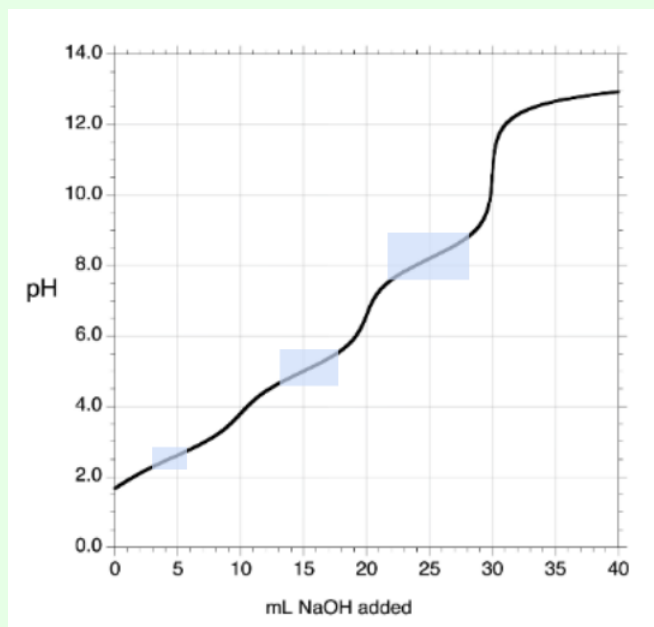
In addition, some amino acids are **triprotic** or **polyprotic**, with three or more  $pK_a$  values. This occurs as a result of the variability of amino acids—amino acids are characterized by a carboxylic acid group, an amine group, and a variable **R group**. The R group may be able to give and take up one or multiple  $H^+$  ions, leading to triprotic or polyprotic amino acids. When the R group cannot do this, the amino acid is always diprotic. The concept of variable R groups is explored more thoroughly in the upcoming section.

**Example 4.2** (USNCO Local Exam 2021) A triprotic acid is titrated with sodium hydroxide to give the titration curve below. What are the  $pK_a$  values of the three acidic hydrogens?



- (A) 2.6, 5.0, 8.2
- (B) 3.8, 6.6, 10.6
- (C) 3.8, 8.2, 12.7
- (D) 5.0, 8.2, 12.7

**Solution:** Recall that buffer regions are flat sections of a titration curve where a buffer is effective. We expect to see three buffer regions on the graph, due to the titrated solution containing a triprotic acid. We can see three buffer regions (look for decreases in slope).



Buffer regions highlighted in blue

Buffer regions are located around the half-equivalence point, located at where  $\text{pH} = \text{p}K_a$ . By looking for the midpoint of each buffer region, we can determine the  $\text{p}K_a$  values. The buffer regions are centered around the points of the graph where  $\text{pH} = 2.6$ ,  $5.0$ , and  $8.2$ . Since at these points,  $\text{pH} = \text{p}K_a$ , the answer is A.

## 5 Biological Macromolecules

In the field of biology, there are four main classes of **macromolecules** important for life. Macromolecules are molecules with up to thousands and thousands of covalently bonded atoms, and these molecules are what are used to carry out life's various processes.

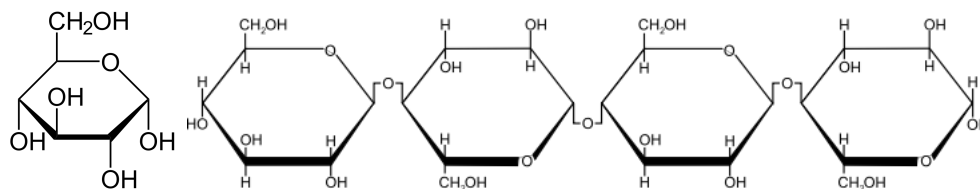
Macromolecules are made up of small building blocks, known as **monomers**. There are different monomers for each type of macromolecule. Monomers are connected through **covalent bonding** (as opposed to the intermolecular force of hydrogen bonding between multiple water molecules), meaning monomers bond together to form one large molecule, known as a **polymer**.

### 5.1 Carbohydrates

**Carbohydrates** are sugars and starches.

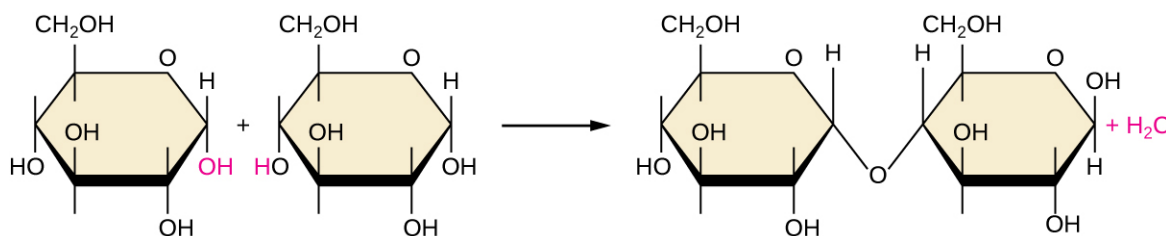
- The monomers of carbohydrates are known as **monosaccharides**. Some examples of monosaccharides include the simplest sugars, like glucose, fructose, and galactose.

- The polymers of carbohydrates are known as **polysaccharides**. Polysaccharides are made of covalently bonded monosaccharides, and include more complex starches such as cellulose. Other common polysaccharides include starch, amylopectin, and amylose,
- **Disaccharides** are also common, made of two monosaccharides. A common disaccharide is sucrose, made of a molecule of glucose and fructose covalently bonded to each other. Other common disaccharides include maltose and lactose.



**Figure 5.1** Glucose, a monosaccharide (left), and cellulose, a polysaccharide (right).

Carbohydrates are useful due to the energy that their bonds store. In order to bond monomers together, energy must be added to break covalent bonds already present inside the molecule. This is necessary, as currently the molecule is stable and will not form new bonds. Removing atoms removes their shared electrons from the molecule, resulting in an unstable compound which is more likely to bond due to a lack of filled valence electron shells. This is the basis for **dehydration synthesis**, energy is added to remove atoms (two hydrogen atoms and one oxygen) from two molecules, making them bond to each other to fill their valence electron shells.



**Figure 5.2** H<sub>2</sub>O being removed from two glucose molecules, resulting in a disaccharide and a water molecule. (Source: CNX OpenStax)

Since dehydration reactions require energy to create larger molecules, the opposite is true. In **hydrolysis**, water is added to a molecule to split it into smaller components, resulting in the release of energy. This is important as it allows carbohydrates and other polymers to act as a form of energy storage—sugars can be used to store energy in the form of bonds, then released when necessary to provide that energy again. For example, potato plants store their glucose molecules in the form of starch, a complex polysaccharide.

Carbohydrates are also used as structural polymers, for example cellulose makes up a significant amount of cell wall composition in many trees.

## 5.2 Lipids

Lipids are neither true polymers nor true macromolecules. They cannot be connected together to create long polymer chains, nor are they large enough to be classified as a macromolecule, however they are still very important biologically. Lipids are a very diverse group of molecules, however

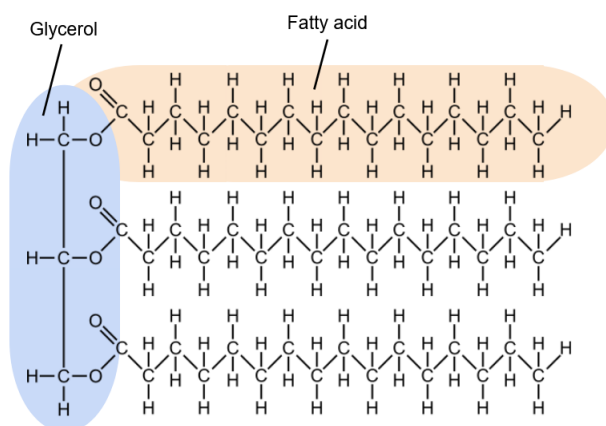


they all have one thing in common, they are **hydrophobic**. Many others are also **amphipathic**, meaning they have both hydrophobic and hydrophilic regions.

There are three important types of lipids to biological life, **fats**, **phospholipids**, and **steroids**.

### 5.2.1 Triglycerides

**Triacylglycerol** is an important type of fat classified under **triglycerides**, made up of one molecule of **glycerol** serving as a backbone, and three molecules of fatty acids connected to it.

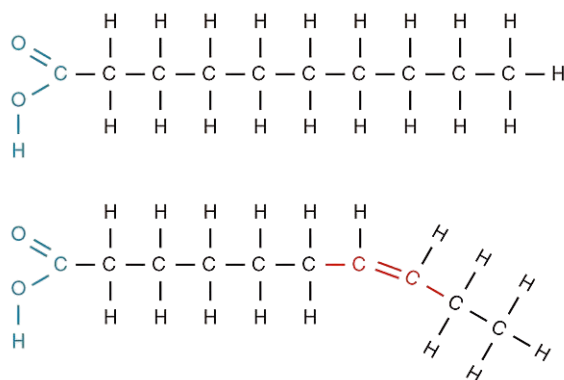


**Figure 5.3** A diagram depicting the structure of a fat.

The fatty acids are nonpolar, as the differences in electronegativity between the hydrogen and carbon bonds which make up the molecule are not great enough to cause significant partial charges. This means that it does not interact with water molecules, there are no electrical charges in the fatty acids to form intermolecular forces between it and the polar water molecule. This is what gives lipids their hydrophobic properties.

Not all fats share the exact same chemical composition. In the above image of a fat, every carbon (except the ends) in the fatty acids is attached to two hydrogen atoms and two carbon atoms, sharing four electrons and completing its valence electron shell. This type of fat is known as a **saturated fat**, every carbon is attached to as many hydrogens as possible.

In some other fats, this is not the case. A carbon may only share electrons with one hydrogen and two other carbons, sharing an extra electron with another carbon atom to reach four shared electrons and a completed valence electron shell. This creates a covalent double bond (due to *two electrons* being shared), causing a **kink**, or crookedness, in the fatty acid. This kink greatly changes the shape of the resulting fat, and when packed together with other fat molecules, cannot be packed closely enough to form a solid. Instead, these **unsaturated fats** are liquids at room temperature, and are commonly referred to as **oils**. The carbons in the fatty acid are not bonded to the maximum amount of hydrogen atoms possible, and it is thus known as unsaturated.



**Figure 5.4** A diagram depicting a saturated fatty acid tail (top) and an unsaturated fatty acid tail (bottom). Note the presence of a double bond and "missing" hydrogen atoms from the unsaturated fatty acid, and also the resulting shape change. (Source: OpenStax College)

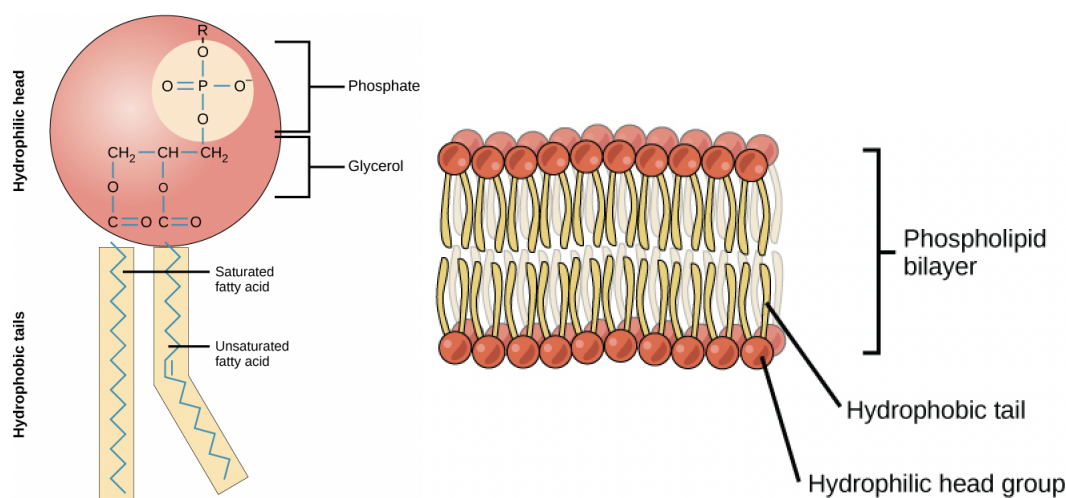
Fats can store a significant amount of densely-packed energy due their large amount of bonds with stored energy. They are used as energy storage by both animals and plants, for example in the form of fat-storing adipose tissue in humans and vegetable oils in plants.

## 5.2.2 Phospholipids

**Phospholipids** are another type of important lipids, and are **amphipathic**. Recall that amphipathic molecules have both hydrophobic and hydrophilic regions. In a phospholipid, there is a glycerol backbone and only two fatty acids attached. Instead of a third fatty acid, a negatively-charged **phosphate** is attached, and commonly another molecule is also attached to the phosphate. This molecule can be one of many, meaning there are a great variety of phospholipids which can be synthesized, each with their own properties.

Since the phosphate head and attached molecule are electrically charged, they can interact with the partial charges of water molecules. This is what causes their hydrophilic properties.

When a group of phospholipids are introduced to a solution of water, they form a **phospholipid bilayer**, a double layer of phospholipids, with each layers' phospholipid heads pointing outwards towards the water, and their tails pointing inwards towards each other.



**Figure 5.5** The basic structure of a phospholipid (left), and a phospholipid bilayer (right). Note how the heads point outward and the tails point inward; this is due to the hydrophilic and hydrophobic properties of the phospholipid. (Source: OpenStax Biology)

This property of phospholipids to form a bilayer is the central basis of all lipid cell membranes. Phospholipid bilayers make up the plasma membrane encapsulating a cell's cytosol, various organelles, and more, making cellular life and compartmentalization possible.

**Example 5.1** (USABO Open Exam 2008) The fluidity of a lipid bilayer is enhanced with:

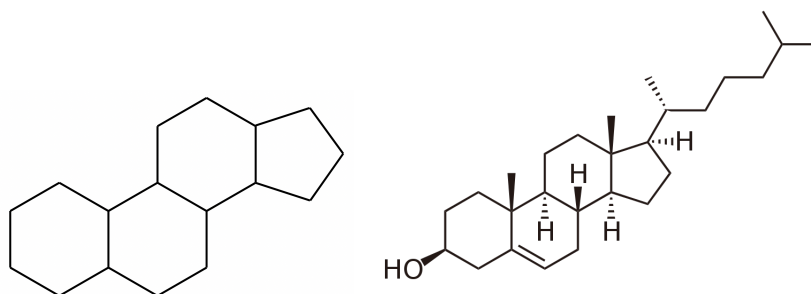
- (A) Decreased temperature
- (B) Increased unsaturation in fatty acid tails
- (C) Increased unsaturation in polar head groups
- (D) Increased saturation in fatty acid tails
- (E) Increased fatty acid chain length

**Solution:** Recall that unsaturated fatty acid tails contain kinks, decreasing their ability to be closely packed. This is the reason why unsaturated fats are liquids at room temperature—they cannot pack closely enough to form a solid. Thus, increasing unsaturation in fatty acid tails would increase the fluidity of a lipid bilayer, so the answer is B.

### 5.2.3 Steroids

**Steroids** are a third group of molecules classified under lipids. They consist of four interconnected carbon rings (multiple carbon atoms bonded together forming a ring). Steroid lipids can vary through what atoms are attached to these rings.

**Cholesterol** is an example of a steroid, commonly incorporated into the plasma membrane along with phospholipids. Other steroid lipids include hormones such as testosterone, estrogen, and cortisol, which regulate important physiological functions, such as growth, reproduction, and metabolism.



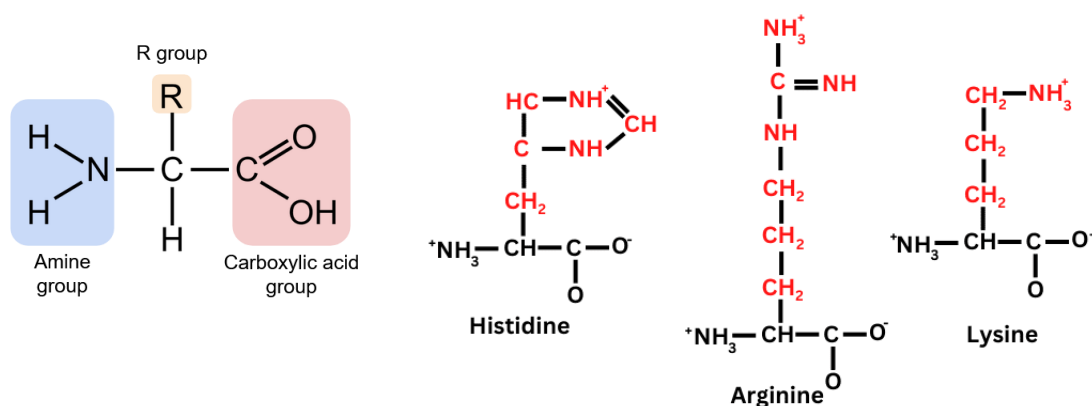
**Figure 5.6** The basic structure of a steroid (left), and the structure of cholesterol (right). Note the four fused rings, each joint in the ring represents a carbon atom.

### 5.3 Proteins

**Proteins** are macromolecules which carry out and make possible many cellular processes, acting as enzymes, transporters, and more.

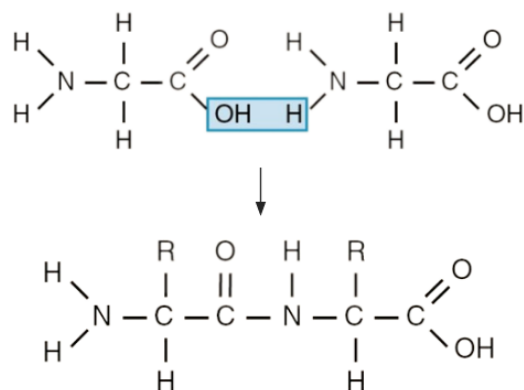
The monomers of proteins are the previously mentioned **amino acids**. Remember that amino acids contain a **carboxylic acid group**, an **amine group**, and a variable **R group**. This R group, also known as a **side chain**, can contain many different atoms and elements, and are what differentiate each amino acid from each other. The R group of the amino acid **glycine** is simply a hydrogen atom, meanwhile the R group of the amino acid arginine contains multiple atoms in a guanidino group (this group is able to give and take  $H^+$  ions, which is what makes arginine triprotic as opposite to diprotic).

The R groups are what determine an amino acid's properties: buffering capabilities, hydrophobic properties, polarity, compactibility, and more. There are twenty unique R groups, and thus twenty unique amino acids, each with their own properties.



**Figure 5.7** The basic structure of an amino acid (left), and the structures of three amino acids (right). Note the differing R groups between histidine, arginine, and lysine (highlighted red). (Source: ScienceQuery)

Amino acids are covalently bonded together through **dehydration synthesis**.  $H_2O$  is formed through removing atoms from one amino acid's amine group, and another amino acid's carboxylic acid. The amino acids then covalently bond together.



**Figure 5.8** Bonding of two amino acids via dehydration synthesis. (Source: OpenStax College)

Polymers formed from amino acid monomers are known as **polypeptides**, or **proteins**. These polypeptides can have many different amino acids, which affect its resulting protein's form and function. Polypeptide chains fold into proteins in a process called **protein folding**, the details of which are discussed in a different handout (see part two). The folding process is determined by the amino acids in a polypeptide—hydrophobic amino acids and hydrophilic amino acids can clump together and act similarly to a lipid bilayer, some amino acids may be attracted to others due to their polarities, some amino acids change depending on the pH of a solution.

Protein structures can thus vary greatly, and this is the reason why we see so many proteins with so many wildly differing functions. Without proteins, life would simply not be possible due to how many vital processes they maintain and allow to exist.

## 5.4 Nucleic Acids

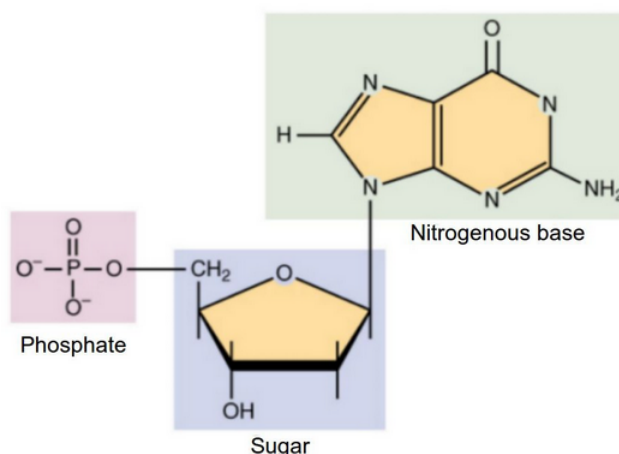
**Nucleic acids** are the basis of our genetic code, and are what allow the phenomenons of evolution and heredity to occur.

The monomers of nucleic acids are known as **nucleotides**. Nucleotides are made up of three main parts, a carbohydrate sugar with five carbon atoms, known as a **pentose**, a nitrogen-containing base, known as a **nitrogenous base**, and a **phosphate group**.

Nitrogenous bases come in two types:

- **Pyrimidines**, which have a six-atom ring of carbon and nitrogen. Pyrimidines include **cytosine (C)**, **thymine (T)**, and **uracil (U)**.
- **Purines**, which have a six-atom ring and a five-atom ring fused together, made of carbon and nitrogen. Purines include **adenine (A)** and **guanine (G)**.

The types of purines and pyrimidines differ in atoms and chemical groups attached to their rings.



**Figure 5.9** The structure of a nucleotide. (Source: OpenStax College)

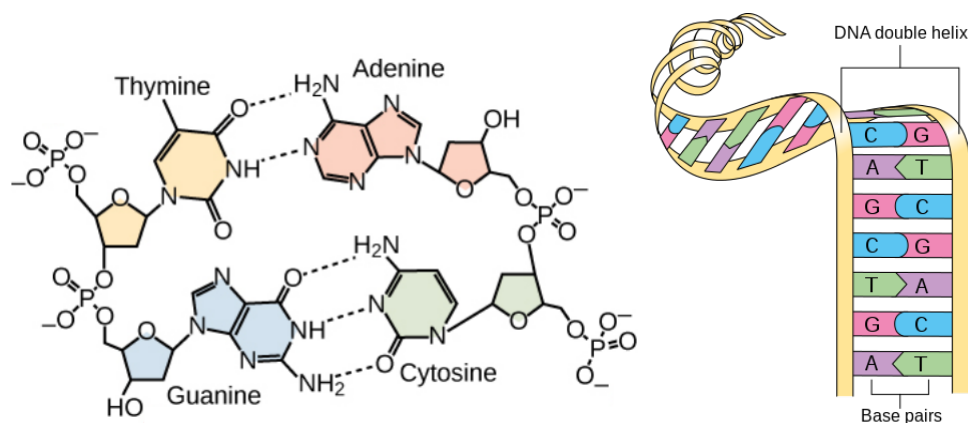
The polymers made from nucleic acid monomers are known as **polynucleotides**, or **nucleic acids**. Some notable nucleic acids include **deoxyribonucleic acid (DNA)** and **ribonucleic acid (RNA)**.

Nucleotides are joined together by **dehydration synthesis**, removing  $H_2O$  from atoms from one nucleotide's pentose sugar, and another nucleotide's phosphate. Nucleotides are thus joined in a chain and supported by what is known as the **sugar phosphate backbone**. Since only the sugar and phosphate are involved in a nucleic acid's structural integrity, the nitrogenous base is free to bond with other molecules. The covalent bonds formed in the creation of the sugar phosphate backbone are known as **phosphodiester bonds** and are much stronger than the relatively weak hydrogen bonds connecting pairs of nitrogenous bases together.

In **deoxyribonucleic acid (DNA)**, the pentose sugar is a molecule known as **deoxyribose**. DNA comes in the shape of a **double helix**, of two winded polynucleotide strands. The two strands are connected by their nitrogenous bases, in a process called **base pairing**. It is important to note that the two strands are **antiparallel**, and essentially that one strand is upside-down compared to the other.

In base pairing, hydrogen bonds are formed between two bases. In DNA, only **adenine and thymine** and **cytosine and guanine** can form hydrogen bonds. These two pairs' chemical compositions are specific enough and complement each other to create suitable hydrogen bonds.

In RNA, the nitrogenous base **uracil** effectively replaces **thymine**, and the pentose sugar is **ribose** instead of deoxyribose.



**Figure 5.10** Diagram showing paired strands of DNA through hydrogen bonding (dashed lines) and base pairing (left), and a model of the double helix structure of DNA (right). (Source: UK Cancer Research and OpenStax Biology)

**Example 5.2** (USABO Open Exam 2003) When the base composition of DNA from bacterium *Mycobacterium tuberculosis* was determined, 18 percent of the bases were found to be adenine. What is the combined guanine and cytosine content in *M. tuberculosis*?

- (A) 18 percent
- (B) 32 percent
- (C) 36 percent
- (D) 64 percent

**Solution:** Recall that adenine can only pair with thymine, and guanine can only pair with cytosine. Since 18 percent of the bases in the DNA were adenine, we know that 18 percent of the bases are thymine.  $18 + 18 = 36$ , so the combined adenine and thymine content of the DNA is 36 percent. The only other nitrogenous bases found in DNA are guanine and cytosine, so we know that the rest of the DNA is made of only guanine and cytosine.  $100 - 36 = 64$ , so the answer is D.

The order of nitrogenous bases in DNA is the basis of our **genetic code**. The order of DNA is transcribed into RNA, which then acts as instructions for amino acid and protein synthesis. Each amino acid is represented by one or more **codons**, or triplets of RNA or DNA nucleotides. For example, one of the codons which encodes for the amino acid glycine is GGC, representing a sequence of guanine, guanine, and cytosine in a strand of RNA, which was transcribed from DNA. DNA thus encodes for proteins, and proteins are what make many cellular processes possible.

Without DNA, there would be no instructions for what proteins to make or how to make them, and effectively nothing would get done. DNA is our genetic code, which determines how everything is made.

## 6 Conclusion

### 6.1 Closing Statement

As you venture through the various topics of biology, you will soon begin to realize how interconnected life is. Every concept, molecule, law, all comes together to create the phenomenon of life. The basic structure of an atom leads to complex molecules. The concept of acidity, based off of said atoms and complex molecules, greatly affects protein folding and function. These proteins which carry out so many processes, like the breakdown of carbohydrates and fats for energy, are made up of amino acid monomers. These amino acids are encoded by nucleotides, which form massive strands of DNA. All of this happens in solutions of water molecules, with their own unique properties which make life possible. There are still many more concepts to grasp, which all play their own part in making life possible.