UNIT 1: METABOLIC PROCESSES

The Biochemical Basis of Life

It is imperative that we understand the concepts in chemistry that are relevant to the study of biology so that we can make sense of the processes that take place in living organisms

Basic Chemical Principles

ATOMIC VARIETIES

- each atom has its own specific number that represents how many protons it, and only it, possesses
- this number is called the **atomic number**
- the number of neutrons in any atom is commonly the same as the number of protons
- the **atomic mass** of any atom is the sum of the number of protons plus the number of neutrons
- some atoms exist in more than one “variety” or “flavour”
- for example, hydrogen is found to exist in three main forms in the universe:
  - *normal hydrogen* – 99.9% of all hydrogen
  - *deuterium* – 0.2% of the rest
  - *tritium* – exists in trace amounts
- notice that all three “varieties” of hydrogen possess 1 proton (i.e. they all have the same atomic number), yet they each possess a different atomic mass, which makes them all **isotopes**
- all isotopes possess similar chemical properties but different physical properties - namely mass
- for example, the element carbon, C, a major constituent of living organisms, consists of three isotopes:
  - *carbon-12* - accounts for 99% of the carbon atoms in nature
  - *cabon-13* – accounts for the rest
  - *carbon 14* - exists in trace amounts
- isotopes naturally break down (decay) thereby releasing subatomic particles and radiation – such isotopes are called radioisotopes
- of the three varieties of carbon, carbon-14 is radioactive
- Table 1, p. 8 summarizes the basic characteristics of the three isotopes of carbon and hydrogen
- all radioactive isotopes have a characteristic half-life – the time it takes for one half of the atoms in a sample to decay
- the half-life rate is constant for each isotope
- two useful applications of radioisotopes are:

1. ***radiometric dating***

- radioactive carbon-14 becomes incorporated into living tissue
- when the tissue dies and becomes decomposed, the carbon-14 that it possesses begins to decay at a constant half-life rate
- when the ratio of carbon-12 to carbon-14 in a dead or fossilized organism is measured, scientists can predict the amount of time that has elapsed since the organism's death

2. ***radioactive tracers***

- when radioactive elements exist in living tissue, they emit radiation
- this radiation can be detected using various kinds of equipment – which means that any radioactive element can be followed or traced chemical reactions
- this is how scientists learn about reaction mechanisms and biochemical processes such as respiration and photosynthesis
- carbon-14 and hydrogen-3 (tritium) are commonly used tracers in biological research
- radioisotopes are also used in diagnoses and treatment
- for example, the thyroid gland, an important organ that regulates human metabolic activity and growth, actively absorbs iodine
- doctors are able to inject radioactive iodine into patients that possess abnormal thyroid activity to help diagnose their thyroid condition – normal, enlarged (over active), or cancerous
- the radioactivity is detected with a photographic device, creating an image of the thyroid gland (Figure 2, p. 9)
- Table 2, p. 10 shows commonly used radioisotopes in nuclear medicine diagnoses

Homework: Practice 1-7, p. 10

BONDING

- **molecules** are the smallest units of a substance that still possess the fundamental chemical and physical properties of the substance
- molecules can be chemically broken down into simpler constituents called **atoms**
- atoms have much different properties when they're isolated than when they're components of a molecule
- **elements** are substances that consist of atoms of only one kind
- living matter tends to have common elements in it – C, H, O, N, P, and S

- it was once thought that atoms could not be broken down into smaller components, hence the name “atom” meaning “indivisible”
- however, today it is known that atoms can be broken down into their sub-atomic constituents: **electrons, protons, and neutrons**
- electrons are negative in charge and possess negligible mass
- protons are positive in charge and possess approximately the same mass as neutrons, which have no charge
- protons and neutrons combine to form the **nucleus** – which is 99.9% of the mass of the atom, but occupies less than 1% of its volume
- using statistics, scientists can determine the most probable location of electrons in regions of space called orbitals
- they are able to determine regions of space where they are most likely to exist
- these fixed, 3-dimensional, regions of space around the nucleus are called **orbitals** (Figure 4, p. 11)
- orbitals can only accommodate 2 electrons
- each energy level that surrounds a nucleus of an atom possesses subshells that contain these orbitals
- for example, energy level one possesses one subshell, (the s subshell), which in turn, is the first orbital, energy level two possesses two subshells, (the s and the p subshell), therefore 4 orbitals, the s, and the three p orbitals, energy level three possesses three subshells, (the s, the p, and the d subshell), therefore nine orbitals, the one s, the three ps, and the five ds, etc.
- the 1st orbital of every energy level has the same shape, the 2nd orbital has another distinct shape (see Figure 4, p. 11)
- the maximum number of electrons that each energy level can hold can be calculated using $2n^2$, where n is the energy level
- for example, energy level 3, alone, can hold a maximum of $2(3)^2 = 18$ electrons
- an atom that has three energy levels can hold a maximum of $2(1)^2 + 2(2)^2 + 2(3)^2 = 28$ electrons
- the arrangement of electrons in the orbitals is called an atom’s electron configuration (Table 3, p. 12)
- the outermost orbitals contain the electrons furthest away from the nucleus of an atom
- the orbitals that exist on the outer-most level contain the electrons that are responsible for the interaction of atoms to form molecules
- electrons found in these outer-most orbitals are called **valence electrons**
- these electrons are called valence electrons
- they are the ones involved in the chemical reactions of that atom
- the chemical stability of an atom is determined by the arrangement of an atom’s valence electrons

- atoms that have completely filled orbitals are more stable, and less reactive than atoms with half-filled, or incomplete orbitals
- all the elements in group 18 of the periodic table have a full set of electrons in their valence shell, therefore they are chemically stable
- all other elements in the universe have incomplete outer orbitals, therefore are reactive
- the most reactive elements are those that have one or two more than the full amount, or one or two less than the full amount
- these are found in the first and second column (group), and the sixteenth and seventeenth column (group)
- Figure 5, p. 12 shows the number of valence electrons that the first 20 elements of the periodic table each possess
- notice that the elements in each column possess the same number of valence electrons
- elements can become chemically stable by either taking, losing, or sharing valence electrons
- the elements on the left of the periodic table will lose the appropriate number of electrons to elements of the right side of the periodic table (excluding the last column of course) so that they will possess the full set number
- for example, if a sodium atom was in contact with a chlorine atom, the sodium would lose one electron to the chlorine, resulting in a stable number of 10 electrons (just like neon)
- consequently, the chlorine will pick up one electron and have a stable number of 28 electrons (just like argon)
- as a result, sodium becomes a **cation** with a positive one in charge, and chlorine becomes an **anion** with a negative one charge
- positive sodium is attracted to negative chloride, resulting in a force of attraction that keeps them together called an **ionic bond**
- the compound that results is sodium chloride – an **ionic compound** where each atom is chemically stable
- the force that keeps oppositely charged ions together is called an **electrostatic force of attraction** - it is not a true molecular bond
- molecular forces of attraction are forces that result from the overlapping of valence orbitals (sharing of electrons) between two atoms
- for example, if a carbon atom were in contact with two oxygen atoms, neither would lose or gain electrons
- instead, the carbon would share two electrons with one oxygen, and two with the other
- the region of space where the sharing of electrons takes place is called an intramolecular bond, also known as a **covalent bond**
- groups of atoms held together by covalent bonds are called true molecules -- Table 4, p. 13, shows examples of different compounds
- the forces that hold atoms together in a compound are called **intramolecular forces of attraction**

Homework: Practice 8-9, p. 16

POLARITY DUE TO ELECTRONEGATIVITY

- all the atoms of the periodic table have a certain ability to attract electrons of other atoms – this ability is called **electronegativity**
- atoms on the right upper hand corner of the periodic table are the smallest, and as a result, their positive proton can get close to electrons of other atoms to attract them away from the other atom and bring them over to themselves – this means that these atoms have a high electronegativity
- atoms on the lower left hand corner of the periodic table are the largest, therefore have a low electronegativity
- when two or more atoms combine, the greater their difference in electronegativity, the greater the polarity of that substance
- in all cases of ionic bonding, and in some cases of covalent bonding where sharing of the electron pair is not equal, the molecule results in being **polar** - it has a positive end and a negative end
- this is because the electrons spend more time around one species (the more electronegative one), and less time around another (the less electronegative one)
- this means that each end of the molecule is oppositely charged – one end is slightly positive, the other, slightly negative
- to determine the amount of polarity in a molecule, the electronegativity values of the atoms involved are subtracted from one another
- if the difference is less than 1.7, the molecule is said to be a polar covalent substance
- if the difference in electronegativity greater than 1.7, the molecule is said to be ionic in character (see Figure 8, p. 14)
- for example, hydrogen chloride is more polar than chlorine gas because the difference in electronegativity between hydrogen and chlorine is 2.9 – 2.1 = 0.8, and the difference between the two atoms of chlorine in chlorine gas is 2.9 – 2.9 = 0.0.
- hydrogen chloride is slightly polar, and chlorine gas is completely non-polar (the truest molecule you can get)

THE SHAPES OF MOLECULES

- a molecule's biological function is determined by the physical three-dimensional shape that it possesses
- the types of atoms involved, each with their own number of valence electrons, determine the kinds of bonds that exist between them
- the electron pairs that exist in the molecular (covalent) bonds between atoms dictate the shape of a molecule
- Figure 9, p. 15 shows an example

- a Canadian chemist (R. Gillespie) developed a theory called the valence shell electron pair repulsion theory (VSEPR) to help determine the shape of any particular molecule
- Table 5, p. 15 shows various molecular shapes that some basic, common molecules can possess

INTERMOLECULAR BONDS

- the polarity of an entire molecule is dependent on two things – the bond polarity and the molecular shape
- symmetrical molecules (like Figure 10 (a)) are non-polar, while asymmetrical molecules are polar in nature
- all molecules attract other molecules – these forces of attraction are called **intermolecular bonds**
- these are the bonds that are broken in a substance when it changes state from solid to liquid to gas
- there are three types of intermolecular bonds, or van der Waals Forces: (Figure 12, p. 17)

1. **London forces** – weakest of the three; exist between all atoms and molecules; occur between non-polar substances

2. **dipole-dipole forces** – hold polar molecules together; positive side of one molecule with the negative end of another

3. **hydrogen bonds –** strongest of the three; occur between a hydrogen of one molecule and a very electronegative atom of

another neighboring molecule, such as nitrogen (N), oxygen (O), or fluorine (F)

- Figure 13 and 14, p. 18 shows a diagram of H-bonding

WATER – THE UNIVERSAL SOLVENT

- water is a very important biological molecule – it is found in large percentages in all living forms
- it is a polar covalent molecule, where the two hydrogens bond with the central oxygen, creating an angle of 104.5 (CHUM FM).
- this shape gives water its polar nature
- the polar nature of water causes intermolecular bonds
- water is considered a universal solvent – more substances dissolve in water than in any other substance
- the reason for this is because of its unique polarity – it has partial positive and partial negative to provide attachment with other molecules (see Figure 15, p. 18)
- all ionic substances dissolve in water and any polar covalent substance dissolves in water
- this is because "like dissolves like", meaning polar substances are **miscible** in other polar substances, and non-polar substances are miscible in other non-polar substances
- for example, water and oil don't mix, because water is polar and oil is non-polar
- thus water and oil are **immiscible**
- water and vinegar mix because both are polar substances
- oxygen does not dissolve in water that well (or blood, since blood is mostly water)
- that is why hemoglobin (a carrier molecule of oxygen) is necessary in blood – it increases the amount of O_2 that can dissolve in blood
- large, non-polar molecules, such as fats and oils are considered **hydrophobic** (meaning "water-fearing") since they cannot form hydrogen bonds with water
- polar molecules are **hydrophilic** (meaning "water-liking") since they can form hydrogen bonds with water
- molecules that contain both hydrophobic and hydrophilic parts are called **amphophilic** molecules

WATER'S UNIQUENESS

- water's angular shape and hydgrogen-bonding characteristics give it extra-ordinary properties
- the following table summarizes, explains, describes the effects, and gives an example of each unique property of water:

What Water Does	Property	Explanation	Result	Example
water clings	cohesion	hydrogen bonds form between water molecules	Great surface tension	a tooth pick floats on water
	adhesion	hydrogen bonds form between water molecules and other polar materials	capillary action	water climbs up xylem of trees
water holds onto heat	relatively high specific heat capacity	hydrogen bonding causes water to take in large amounts of heat before its temperature is increases and also causes it to lose large amounts of heat before its temperature decreases significantly	maintenance of temperature	high heat capacity helps organisms maintain a constant body temperature
	high specific heat of vaporization	hydrogen bonding causes liquid water to absorb a large amount of heat to become a vapour (gas)	evaporative cooling	many organisms, including humans, lose body heat by evaporation of water from surfaces, such as skin (by sweating) and tongue (by panting)
solid water is less dense than liquid water	highest density at 4˚C	as water molecules cool below 0˚C, they form a crystalline lattice (freezing) – the hydrogen bonds between the V-shaped molecules spread the molecules apart, reducing the density below that of liquid water	ice floats on liquid water	fish and other aquatic organisms are able to survive in winter

ACIDS AND BASES

- an acid is a substance that possesses **hydronium ions** – H_3O^+
- an acid is sour, conducts electricity, reacts with metals to produce hydrogen gas, turns blue litmus paper red

- a base is a substance that possesses **hydroxide ions** – OH^-
- a base is bitter, has a slippery feel, conducts electricity, and changes red litmus paper blue
- water is equally acidic and basic – it is considered **neutral**
- it breaks up in a manner such that produces exactly equal amounts of each ion: $H_2O\ (l) \leftrightarrow H_3O^+\ (aq) + OH^-\ (aq)$
- when more hydronium exists in solution, the substance is acidic
- for example: $HCl\ (g) + H_2O\ (l) \leftrightarrow H_3O^+\ (aq) + Cl^-\ (aq)$
- when more hydroxide exists in solution, the substance is basic
- for example: $NaOH\ (s) + H_2O\ (l) \leftrightarrow Na^+\ (aq) + OH^-\ (aq)$
- when an acid is mixed with a base, two substances are always produced: a salt and water
- this is called a neutralization reaction
- for example: $NaOH\ (aq) + HCl\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l)$
- to determine the degree of acidity or alkalinity a substance is, the pH scale is used
- solutions with a pH below 7 are acidic – the closer to zero the more acidic
- solutions with a pH above 7 are basic – the closer to 14 the more basic
- solutions with a pH equal to 7 are neutral
- substances that completely ionize to produce hydronium ions are strong acids (i.e. hydrochloric acid)
- substances that barely ionize to produce hydronium ions are weak acids (i.e. vinegar)
- substances that completely ionize to produce hydroxide ions are strong bases (i.e. sodium hydroxide)
- substances that barely ionize to produce hydroxide ions are weak bases (i.e. ammonia)
- according to the Bonsted-Lowry concept, a substance that donates a proton (H^+ ion) is an acid, whereas a substance that accepts a proton is a base
- for example, when acetic acid (vinegar) reacts with water, the following occurs:

$$\overset{H^+}{}$$

$$CH_3COOH\ (aq) + H_2O\ (l) \leftrightarrow CH_3COO^-\ (aq) + H_3O^+\ (aq)$$

$$\underset{H^+}{}$$

- for the above case, since acetic acid donates the proton to the water, and the water accepts it to become hydronium, acetic acid is the acid and water is the base
- in the reverse reaction, since hydronium donates a proton to the acetate, and the acetate accepts it to become acetic acid, hydronium is the acid and acetate is the base
- therefore, acetic acid is the conjugate acid, acetate is the conjugate base, water is the conjugate base, and hydronium is the conjugate acid (see p. 21 of text)

BUFFERS

- most cellular processes operate best at pH 7
- living cells use **buffers** to resist significant changes in pH that could seriously disrupt biological processes
- the most important conjugate acid-base pair buffer is the carbonic – bicarbonate base pair
- for example, if a person's blood gets acidic (has extra H^+ ions), HCO_3^- ions react with the surplus H^+ ions, thereby pulling them out of the blood, and producing H_2CO_3 (aq)
- if a person's blood gets too basic (lacks H^+ ions), H_2CO_3 (aq) dissociates, thereby replenishing the missing H^+ ions
- this buffering effect keeps the blood at a pH of 7.4, which is the ideal pH for internal biochemical processes
- some proteins in the blood can also act as buffers
- for example, hemoglobin has both acidic and basic components to it that can either pull out excess H^+ ions or replenish lost H^+ ions, resulting in a buffering effect in red blood cells

Homework: Section 1.1 Questions 1-15, p. 23