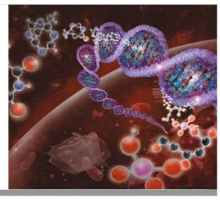
2

The Chemistry of Life



Erin C. Amerman
Florida State College at Jacksonville

Lecture Presentation by Suzanne Pundt
University of Texas at Tyler

Module 2.1 Atoms and Elements

Matter

- **Matter** – anything that has *mass* and *occupies space*; can exist in *three states*: **solid**, **liquid**, or **gas**
- **Chemistry** – study of matter and its interactions

Atoms and Atomic Structure

- **Atom** – smallest unit of matter that retains *original properties*
- Made up of even smaller structures called **subatomic particles**

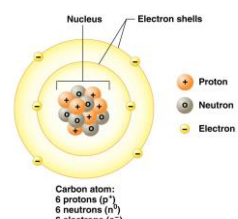


Figure 2.1 Structure of a representative atom.

Atoms and Atomic Structure

- Subatomic particles exist in *3 forms*:
  - **Protons** (**p**$^{+}$) – found in central core of atom (atomic nucleus); *positively* charged
  - **Neutrons** (**n**^{0}) – found in atomic nucleus; slightly larger than protons; *no charge*.
  - **Electrons** (**e**$^{-}$) – found outside atomic nucleus; *negatively* charged
- Atoms are electrically **neutral** – they have *no charge*; number of protons and electrons are equal, cancelling each other's charge; number of *neutrons* does not have to equal number of protons

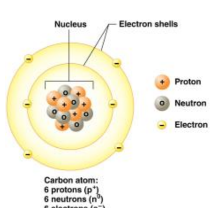
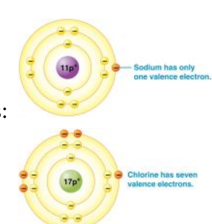


Figure 2.1 Structure of a representative atom.

Atoms and Atomic Structure

- **Electron shells** – *regions* surrounding atomic nucleus where *electrons exist*; each can hold a certain number of electrons:
  - **1st shell** (closest to nucleus) can hold *2 electrons*
  - **2nd shell** can hold *8 electrons*
  - **3rd shell** can hold 18 electrons but "satisfied" with 8
- Some atoms may have more than 3 shells



Elements in the Periodic Table and the Human Body

- Number of **protons** that an atom has in its nucleus is its **atomic number**
- **Atomic number** *defines* every **element**:
  - **Element –** substance that cannot be broken down into simpler substance by *chemical means*
  - Each element is made of atoms with same *number of protons*

Elements in the Periodic Table and the Human Body

- The **periodic table** of elements lists elements by their increasing *atomic numbers*:
  - Organizes elements into *groups* with certain properties
  - Each element is represented by a **chemical symbol**

The Periodic Table

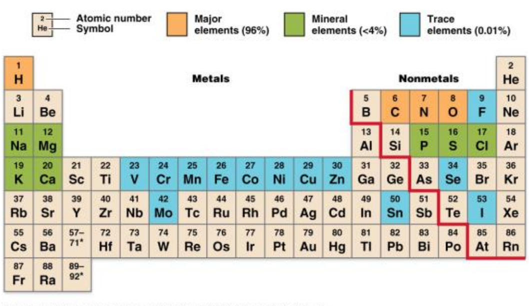


Figure 2.2 Elements in the human body and their positions in the periodic table.

Elements in the Periodic Table and the Human Body

- The human body is made up of *four major* elements:
  - **Hydrogen**
  - **Oxygen**
  - **Carbon**
  - **Nitrogen**
- Also 7 **mineral** elements and 13 **trace** elements

Isotopes and Radioactivity

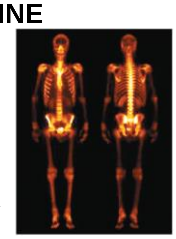
- **Mass number** – equal to *sum of all protons and neutrons* found in atomic nucleus
- **Isotope** – atom with same *atomic number* (same number of protons), but different *mass number* (different number of *neutrons*)

  $^{1}_{1}H$ *Hydrogen* $^{2}_{1}H$ *Deuterium* $^{3}_{1}H$ *Tritium*

- **Radioisotopes** – *unstable* isotopes; high energy or radiation released by *radioactive decay*; allows isotope to assume a more stable form

Nuclear Medicine

Common applications of radioisotopes:



- **Cancer radiation therapy** – radiation *damages* structure of cancer cells; interferes with functions
- **Radiotracers** – injected into patient and detected by camera; image analyzed by computer; shows *size*, *shape*, and *activity* of organs and cells
- **Treatment of thyroid disorders** – high doses of **iodine-131** treat *overactive* or *cancerous* thyroid tissue; radioisotope accumulates and damages cells

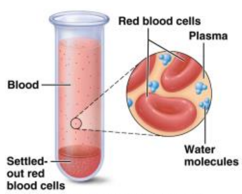
Module 2.2 Matter Combined: Mixtures and Chemical Bonds

Matter Combined

- **Matter** can be combined physically to form a **mixture** – atoms of two or more elements *physically intermixed* without changing *chemical nature* of atoms themselves
- There are 3 basic types of mixtures: **suspensions**, **colloids**, and **solutions** (**Figure 2.3**)

Mixtures

- **Suspension** – mixture containing two or more components with large, *unevenly distributed particles*; will *settle out* when left undisturbed

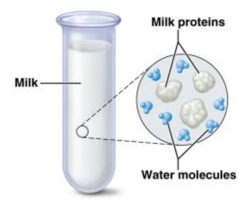


(a) Suspension: Particles large and usually visible; settle out.

Figure 2.3a The three types of mixtures.

Mixtures

- **Colloids** – two or more components with small, *evenly distributed particles*; will not *settle out*

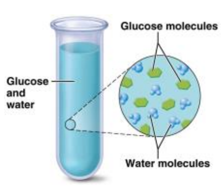


(b) Colloid: Two distinct components; particles small and not visible; do not settle out.

Figure 2.3b The three types of mixtures.

Mixtures

- **Solutions** – two or more components with extremely small, evenly distributed particles; will not *settle out*; contain a **solute dissolved** in a solvent:
  - **Solute** – substance that is *dissolved*
  - **Solvent** – substance that *dissolves* solute

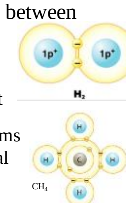


(c) Solution: Particles extremely small and not visible; do not settle out; one component dissolves in the other component.

Figure 2.3c The three types of mixtures.

Chemical Bonds

- Matter can be combined chemically when atoms are combined by **chemical bonds**.
- A chemical bond is not a physical structure but rather an *energy relationship* or *attractive force* between atoms
  - **Molecule** – formed by chemical bonding between two or more atoms of same element
  - **Compound** – formed when two or more atoms from different elements combine by chemical bonding



Chemical Bonds

- **Macromolecules** – very *large molecules* composed of many atoms
- **Molecular formulas** – represent molecules *symbolically* with *letters* and *numbers*; show *kinds* and *numbers* of atoms in a molecule

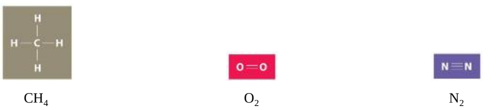
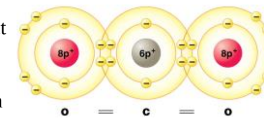


Table 2.1 Electron Sharing in Covalent Bonds.

Chemical Bonds

- Chemical bonds are formed when **valence electrons** (in *outermost* **valence shell**) of atoms interact
- Valence electrons determine *how* an atom interacts with other atoms and whether it will *form bonds* with a specific atom
  - The **octet rule** states that an atom is *most stable* when it has 8 electrons in its *valence shell* (as in CO_2)



The Duet Rule

- The **duet rule** (for atoms with *5 or fewer electrons*) states that an atom is *most stable* when its valence electron shell holds 2 electrons

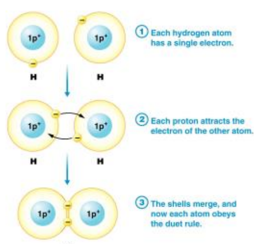


Figure 2.5 Formation of a covalent bond.

Ions and Ionic Bonds

- **Ionic bond** – formed when electrons are *transferred* from a **metal** atom to a **nonmetal** atom; results in formation of **ions**: cations and anions (**Figure 2.4**)
  - **Cation** – *positively charged* ion; forms when metal loses one or more electrons
  - **Anion** – *negatively charged* ion; forms when nonmetal gains one or more electrons
- The attraction between *opposite charges* bonds ions to one another forming a compound called a **salt**

Ionic Bonds

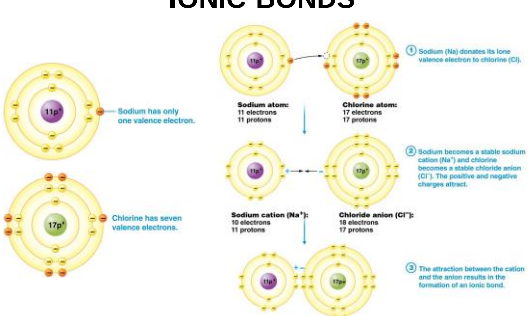


Figure 2.4 Formation of an ionic bond.

Covalent Bonds

- **Covalent bonds** – strongest bond; form when two or more nonmetals *share electrons* (**Figures 2.5**, **2.6**; **Table 2.1**)
- Two atoms can share one (**single bond**), two (**double bond**), or three (**triple bond**) electron pairs:

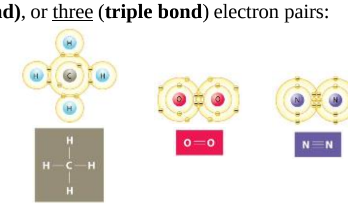


Table 2.1 Electron Sharing in Covalent Bonds.

Covalent Bonds

All elements have protons that *attract electrons*; property known as **electronegativity**:

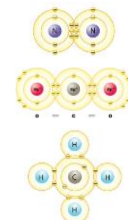
- An element's electronegativity increases from the *bottom left* to the *upper right* of the periodic table making **fluorine** (**F**) the most electronegative element
- The more electronegative an element the more strongly it attracts electrons, *pulling them away* from less electronegative elements

Nonpolar Covalent Bonds

Nonpolar covalent bonds result when two nonmetals in a molecule with similar or identical electronegativities pull with *equal force*; therefore share electrons equally (**Figure 2.6a**)

Nonpolar molecules occur in 3 situations:

- Atoms sharing electrons are *same element*
- *Arrangement of atoms* makes one atom unable to pull more strongly than another atom (as in CO_2)
- Bond is between *carbon* and *hydrogen*



Nonpolar Covalent Bonds

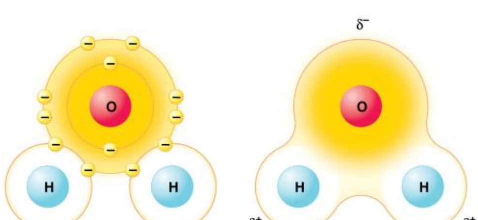
(a) Nonpolar covalent bond—H_2 (hydrogen molecule): Electrons spend equal time around the two hydrogen atoms.

Figure 2.6a Nonpolar vs. polar covalent bonds.

Polar Covalent Bonds

- **Polar** covalent bonds form **polar molecules** when nonmetals with different electronegativities interact resulting in an *unequal sharing* of electrons (**Figure 2.6b**)
  - Atom with higher electronegativity becomes *partially negative* (δ–) as it pulls shared electrons close to itself
  - Atom with lower electronegativity becomes *partially positive* (δ+) as shared electrons are pulled toward other atom
- Polar molecules with *partially positive* and *partially negative ends* are known as **dipoles**

Polar Covalent Bonds



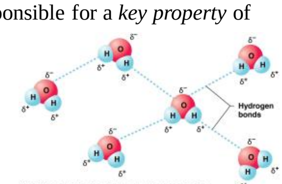
(b) Polar covalent bond—H_2O (water): Electrons spend more time around the more electronegative oxygen atom.

Figure 2.6b Nonpolar vs. polar covalent bonds.

Hydrogen Bonds

Hydrogen bonds – weak attractions between *partially positive* end of one dipole and *partially negative* end of another dipole

- Hydrogen bonds are responsible for a *key property* of water—**surface tension**

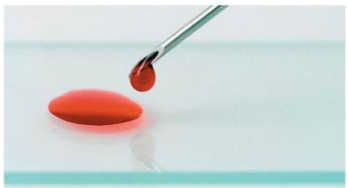


(a) Hydrogen bonds between water molecules

Figure 2.7a Hydrogen bonding and surface tension between water molecules.

Hydrogen Bonds

- Polar water molecules are more strongly attracted to *one another* than they are to nonpolar air molecules at surface



(b) Hydrogen bonds between water molecules create surface tension that causes blood to form droplets.

Figure 2.7b Hydrogen bonding and surface tension between water molecules.

Concept Boost: Determining the Type of Bonds in a Molecule

Basic "rules" to keep in mind:

- If the compound contains both a metal and a nonmetal, the bond is *ionic*
- If the molecule contains two or more nonmetals, the bond is *covalent*; hydrogen behaves like a nonmetal:
  - If the molecule contains two identical nonmetals, it is *non-polar covalent* (e.g., O_2)
  - If the molecule contains only or primarily carbon and hydrogen, it is *nonpolar covalent* (e.g., CH_4)

Concept Boost: Determining the Type of Bonds in a Molecule

Basic "rules" to can keep in mind (continued):

- If the molecule contains two or more nonmetals, the bond is covalent; hydrogen behaves like a nonmetal
  - If the molecule contains two nonmetals of significantly different electronegativities, it is *polar covalent* (hydrogen and carbon have low electronegativities, whereas elements like oxygen, nitrogen, and phosphorus have high electronegativities)

Module 2.3 Chemical Reactions

Chemical Notation

- A **chemical reaction** has occurred every time a chemical bond is *formed*, *broken*, or *rearranged*, or when *electrons are transferred between* two or more atoms (or molecules)
- **Chemical notation** – series of *symbols* and *abbreviations* used to demonstrate what occurs in a reaction; the **chemical equation** (basic form of chemical notation) has two parts:
  - **Reactants** on *left side* of equation are starting ingredients; will undergo *reaction*
  - **Products** on *right side* of equation are *results* of chemical reaction

Chemical Notation

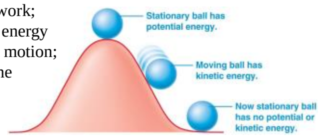
- **Reversible** reactions can proceed in *either direction* as denoted by two arrows that run in opposite directions (as below)
- **Irreversible** reactions proceed from *left to right* as denoted by a single arrow

$$CO_2 + H_2O \leftrightarrows H_2CO_3$$

Reactants (carbon dioxide + water) **Product** (carbonic acid)

Energy and Chemical Reactions

- **Energy** is defined as capacity to do **work** or put matter into *motion* or *fuel* chemical reactions; two *general forms* of energy:
  - **Potential energy** is *stored*; can be released to do work at some later time
  - **Kinetic energy** is *potential energy* that has been released or set in motion to perform work; all atoms have kinetic energy as they are in constant motion; the faster they move the greater that energy



Energy and Chemical Reactions

Energy is found in 3 forms in the human body; chemical, electrical, and mechanical, each of which may be *potential* or *kinetic* depending on location or process

- **Chemical energy** – found in *bonds* between atoms; drives nearly all chemical processes
- **Electrical energy** – generated by movement of *charged particles* or *ions*
- **Mechanical energy** – energy directly transferred from *one object* to *another*

Energy and Chemical Reactions

Energy, inherent in all chemical bonds, must be invested any time a chemical reaction occurs:

- **Endergonic** reactions require *input* of energy from another source; products contain more energy than reactants because energy was invested so reaction could proceed
- **Exergonic** reactions *release* excess energy so products have less energy than reactants

Homeostasis and Types of Chemical Reactions

Three fundamental processes occur in the body to maintain homeostasis (breaking down molecules, converting the energy in food to usable form, and building new molecules); carried out by three basic *types of chemical reactions*:

1. **Catabolic** reactions (**decomposition** reactions)
2. **Exchange** reactions
3. **Anabolic** reactions (**synthesis** reactions)

Homeostasis and Types of Chemical Reactions

- **Catabolic** reactions (**decomposition** reactions) – when a large substance is *broken down* into smaller substances
- General chemical notation for reaction is

$$AB \rightarrow A + B$$

- Usually **exergonic** because chemical bonds are broken

Homeostasis and Types of Chemical Reactions

- **Exchange** reactions occur when one or more atoms from reactants are *exchanged for one another*
- General chemical notation for reaction is

$$AB + CD \rightarrow AD + BC$$

$$HCL + NaOH \rightarrow H_2O + NaCL$$

Homeostasis and Types of Chemical Reactions

- **Oxidation-reduction** reactions (**redox** reactions) – special kind of *exchange reaction*; occur when <u>electrons</u> and <u>energy</u> are exchanged *instead of atoms*
  - Reactant that <u>loses</u> electrons is **oxidized**
  - Reactant that <u>gains</u> electrons is **reduced**
- Redox reactions are usually exergonic reactions capable of *releasing large amounts of energy*

Homeostasis and Types of Chemical Reactions

- **Anabolic** reactions (**synthesis** reactions) occur when small simple subunits and <u>united</u> by chemical bonds to make large *more complex substances*
- General chemical notation for reaction is

$$A + B \rightarrow AB$$

- These reactions are endergonic; <u>fueled</u> by chemical energy

Reaction Rates and Enzymes

- For a reaction to occur atoms must *collide* with <u>enough</u> <u>energy</u> overcome the *repulsion* of their electrons
- This energy required for all chemical reactions is called the **activation energy** (**Ea**)

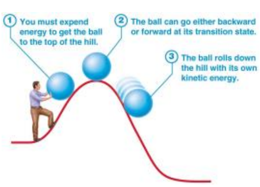


Figure 2.8 Activation energy.

Reaction Rates and Enzymes

Analogy can be applied to chemical reactions – activation energy must be supplied so that reactants reach their *transition states* (i.e., get to the top of the energy "hill") in order to react and form products (i.e., roll down the hill)

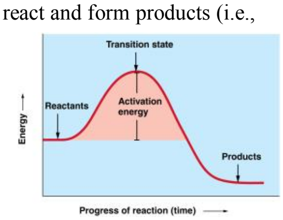


Figure 2.8 Activation energy.

Reaction Rates and Enzymes

- The following factors <u>increase</u> *reaction rate* by <u>reducing</u> *activation energy* or increasing *likelihood of strong collisions* between reactants:
  - Concentration
  - Temperature
  - Reactant properties
  - Presence or absence of a catalyst

Reaction Rates and Enzymes

- When reactant **concentration** increases, more reactant particles are present, increasing chance of *successful collisions* between reactants
- Raising the **temperature** of the reactants increases kinetic energy of their atoms leading to more *forceful* and *effective* collisions between reactants

Reaction Rates and Enzymes

- Both particle **size** and **phase** (solid, liquid, or gas) influence reaction rates:
  - <u>Smaller</u> particles move <u>faster</u> with *more energy* than larger particles
  - Reactant particles in the <u>gaseous</u> phase have *higher kinetic energy* than those in either solid or liquid phase

Reaction Rates and Enzymes

- **Catalyst** – substance that *increases reaction rate* by <u>lowering</u> activation energy without being *consumed* or *altered* in reaction
- **Enzymes** – *biological catalysts*; most are *proteins* with following properties:
  - Speed up reactions by <u>lowering</u> the activation energy (**Figure 2.9**)
  - Highly **specific** for individual **substrates** (substance that can bind to the enzyme's **active site**)
  - <u>Do</u> <u>not</u> alter the reactants or products
  - <u>Not</u> permanently altered in reactions catalyzed

Reaction Rates and Enzymes

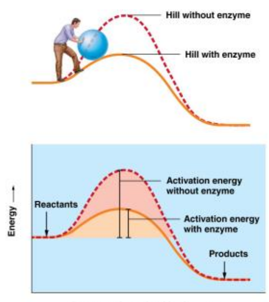


Figure 2.9 The effect of enzymes on activation energy.

Reaction Rates and Enzymes

- **Induced-fit mechanism** – describes enzyme's *interaction* with its substrate(s)
  - Binding of substrate causes a small *shape change* that <u>reduces</u> energy of activation

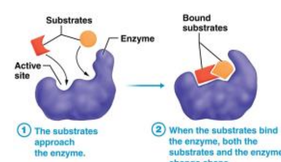


Figure 2.10 Enzyme-substrate interaction.

Reaction Rates and Enzymes

- **Induced-fit mechanism** (continued):
  - Allows transition state to proceed to final products

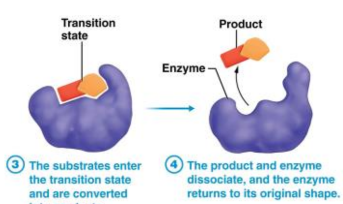


Figure 2.10 Enzyme-substrate interaction.

A&P in the Real World: Enzyme Deficiencies

Examples of *common enzyme deficiencies*:

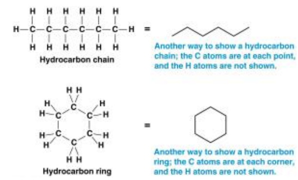
- **Tay-Sachs Disease** – deficiency of **hexosaminidase**; **gangliosides** accumulate around neurons of brain; death usually by age 3
- **Severe Combined Immunodeficiency Syndrome** (**SCIDs**) – may be due to **adenosine deaminase** deficiency; nearly complete absence of immune system; affected patients must live in sterile "bubble"
- **Phenylketonuria** – deficiency of **phenylalanine hydroxylase**; converts **phenylalanine** into **tyrosine**; resulting seizures and mental retardation can be prevented by *dietary modification*

Module 2.4 Inorganic Compounds: Water, Acids, Bases, and Salts Bonds

Biochemistry

Biochemistry – the chemistry of life

- **Inorganic** compounds generally do not contain *carbon bonded to hydrogen*; include **water**, **acids**, **bases**, and **salts**
- **Organic** compounds – those that do contain *carbon bonded to hydrogen*



Water

Water (H_2O) makes up 60–80% of mass of human body and has several *key properties* vital to our existence (**Figure 2.11**):

- High **heat capacity** – able to *absorb heat* without significantly *changing temperature* itself
- *Carries heat with it* when it **evaporates** (when changing from liquid to gas)
- **Cushions** and protects body structures because of relatively high **density**
- Acts as a **lubricant** between two adjacent surfaces (reduces **friction**)

Water

- Water serves as body's *primary solvent*; often called the **universal solvent** because so many solutes will **dissolve** in it entirely or to some degree (**Figure 2.11**)
- Water is a polar covalent molecule:
  - Oxygen pole – *partially negative* (δ^-)
  - Hydrogen pole – partially positive (δ^+)
- Allows water molecules to interact with certain solutes, surround them, and *keep them apart*

Water

- Water is only able to dissolve **hydrophilic** solutes (those with *fully* or *partially charged* ends); "**like dissolves like**", so water dissolves *ionic* and *polar covalent* solutes

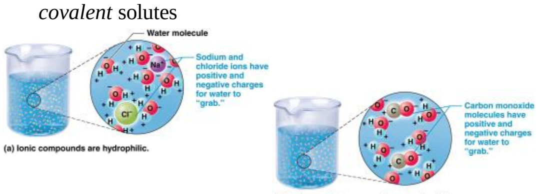


Figure 2.11a, b The behavior of hydrophilic and hydrophobic molecules in water.

Water

- Solutes that do not have full or partially charged ends are **hydrophobic**; do not dissolve in water; includes *uncharged nonpolar covalent* molecules such as **oils** and **fats**

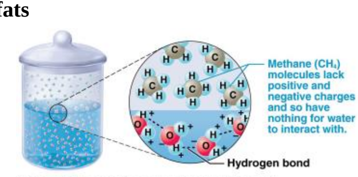


Figure 2.11c The behavior of hydrophilic and hydrophobic molecules in water.

Acids and Bases

- The study of **acids** and **bases** is really the study of the **hydrogen ion** (H^+)
- Water molecules in solution may **dissociate** (break apart) into positively charged **hydrogen ions** (H^+) and negatively charged **hydroxide ions** (OH^-)
- Acids and bases are defined according to their *behavior* with respect to hydrogen ions (next slide)

$$H_2O \rightleftarrows H^+ + OH^-$$

Water — *Hydrogen ion* — *Hydroxide ion*

Figure 2.12a The behavior of acids and bases in water.

Acids and Bases

- **Acid** – hydrogen ion or **proton donor**; number of hydrogen ions <u>increases</u> in water when acid is added (**Figure 2.12b**)
- **Base** (**alkali**) – hydrogen ion <u>acceptor</u>; number of hydrogen ions <u>decreases</u> in water when base is added (**Figure 2.12c**)

Acids and Bases

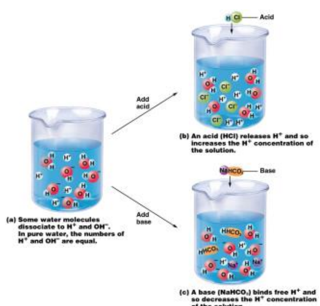


Figure 2.12 The behavior of acids and bases in water.

Acids and Bases

- **pH scale** – ranges from 0–14 (**Figure 2.13**)
- Simple way of representing hydrogen ion **concentration** of a solution
- Literally the *negative logarithm of the hydrogen ion concentration*:

$$\textbf{pH} = -\textbf{Log}\,[\textbf{H}^+]$$

Acids and Bases

- When *pH = 7* the solution is **neutral** where the number of hydrogen ions and base ions are <u>equal</u>
- A solution with *pH less than 7* is **acidic**; hydrogen ions outnumber base ions
- A solution with *pH greater than 7* is **basic** or **alkaline**; base ions outnumber hydrogen ions.

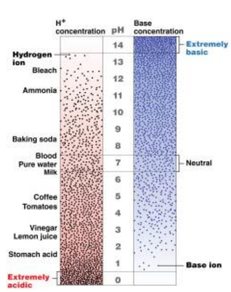


Figure 2.13 The pH Scale.

Acids and Bases

- **Buffer** – chemical system that *resists changes in pH*; prevents large swings in pH when acid or base is *added to a solution*
- Blood pH must remain within its *narrow range* to maintain homeostasis
- Most body fluids are *slightly basic*:
  - Blood pH is 7.35–7.45
  - Intracellular pH is 7.2

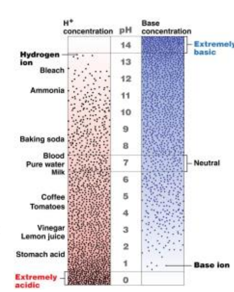


Figure 2.13 The pH Scale.

Concept Boost: Making sense of the pH Scale

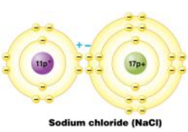
- Why does pH decrease if solution has more hydrogen ions?
- The smaller the pH number, the bigger its negative log
- Single-digit changes in negative logarithm (e.g., from 2 to 3) accompanies a *10-fold change* in hydrogen ion concentration (e.g., from 0.01 to 0.001)

Concept Boost: Making sense of the pH Scale

- Example:
  - Solution A has a hydrogen ion concentration of 0.015 M and a pH of 1.82; solution B has a hydrogen ion concentration of 0.0003 M and a pH of 3.52
  - The solution with the higher hydrogen ion concentration has the lower −log. For this reason, the more acidic a solution, the lower its pH, and vice-versa

Salts and Electrolytes

- **Salt** – any *metal cation* and *nonmetal anion* held together by *ionic bonds*
- Salts can dissolve in water to form cations and anions called **electrolytes** which are capable of *conducting electrical current*



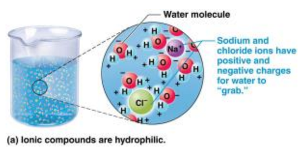


Figure 2.4 and Figure 2.11a

Module 2.5 Organic Compounds: Carbohydrates, Lipids, Proteins, and Nucleotides

Monomers and Polymers

Each type of organic compound in body (**carbohydrate**, **lipid**, **protein**, or **nucleic acid**) consists of polymers built from monomer subunits:

- **Monomers** are *single subunits* that can be combined to build larger structures called **polymers** by **dehydration synthesis** (anabolic reaction that links monomers together and makes a *molecule of water* in process)
- **Hydrolysis** is a catabolic reaction that uses *water to break up polymers* into smaller subunits

Carbohydrates

- **Carbohydrates**, composed of carbon, hydrogen, and oxygen, function primarily as *fuel*; some limited *structural roles*
  - **Monosaccharides** – consist of 3 to 7 carbons; *monomers* from which all carbohydrates are made; glucose, fructose, galactose, ribose, and dexoyribose are most abundant monosaccharides (**Figure 2.14**)

CARBOHYDRATES

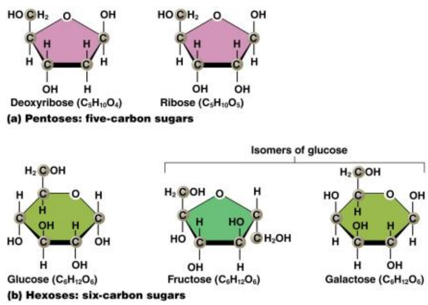


Figure 2.14 Carbohydrates: structure of monosaccharides.

CARBOHYDRATES

- **Disaccharides** are formed by union of <u>two</u> monosaccharides by *dehydration synthesis*

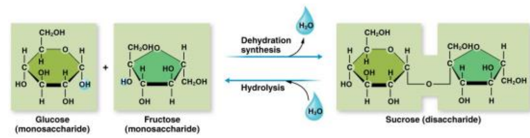


Figure 2.15 Carbohydrates: formation and breakdown of disaccharides.

CARBOHYDRATES

Polysaccharides consist of <u>many</u> monosaccharides joined to one another by dehydration synthesis reactions (**Figure 2.16**)

- **Glycogen** is the *storage* polymer of glucose; mostly in skeletal muscle and liver cells
- Some polysaccharides are found *covalently bound* to either proteins or lipids forming **glycoproteins** and **glycolipids**; various functions in body

CARBOHYDRATES

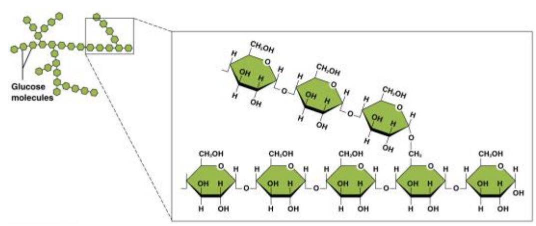


Figure 2.16 Carbohydrates: the polysaccharide glycogen.

LIPIDS

- **Lipids** – group of *nonpolar hydrophobic molecules* composed primarily of carbon and hydrogen; include **fats** and **oils**
- **Fatty acids** – lipid monomers consisting of 4 to 20 carbon atoms; may have none, one, or more double bonds between carbons in *hydrocarbon chain* (**Figure 2.17**)

LIPIDS

- **Saturated** fatty acids – <u>solid</u> at room temperature; have *no double bonds* between carbon atoms so carbons are "saturated" with <u>maximum</u> number of *hydrogen atoms*

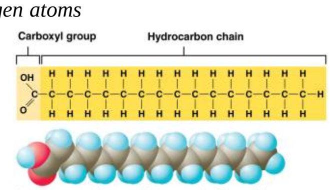


Figure 2.17a Lipids: structure of fatty acids.

Lipids

- **Monounsaturated** fatty acids – generally <u>liquid</u> at room temperature; have <u>one</u> *double bond* between two carbons in hydrocarbon chain

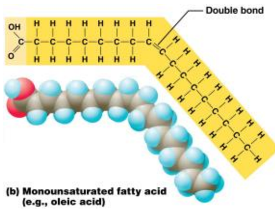


Figure 2.17b Lipids: structure of fatty acids.

Lipids

- **Polyunsaturated** fatty acids – <u>liquid</u> at room temperature; have <u>two or more</u> *double bonds* between carbons in hydrocarbon chain

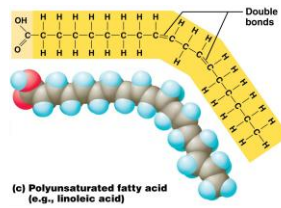


Figure 2.17c Lipids: structure of fatty acids.

The Good, the Bad, and the Ugly of Fatty Acids

Not all fatty acids were created equally:

- *The Good*: **Omega – 3 Fats**
  - Found in flaxseed oil and fish oil but cannot be made by humans; must be obtained in diet
  - Polyunsaturated; positive effects on cardiovascular health
- *The Bad*: **Saturated Fats**
  - Found in animal fats; also in palm and coconut oils
  - Overconsumption associated with increased cardiac disease risk

The Good, the Bad, and the Ugly of Fatty Acids

Not all fatty acids were created equally (continued):

- *The Ugly*: **Trans Fats**
  - Produced by adding H atoms to unsaturated plant oils ("partially hydrogenated oils")
  - No safe consumption level; significantly increase risk of heart disease

Lipids

Triglyceride – three fatty acids linked by dehydration synthesis to a modified 3-carbon carbohydrate, **glycerol**; *storage polymer* for fatty acids (also called a **neutral fat**)

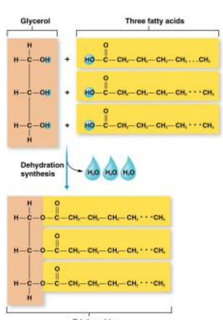


Figure 2.18 Lipids: structure and formation of triglycerides.

Lipids

- **Phospholipids** – composed of a glycerol backbone, <u>two</u> fatty acid "tails" and one *phosphate "head"* in place of third fatty acid (**Figure 2.19**)
- A molecule with a *polar group* (phosphate head) <u>and</u> a *nonpolar group* (fatty acid tail) is called **amphiphilic**
- This amphiphilic nature makes phospholipids <u>vital</u> to the structure of **cell membranes**

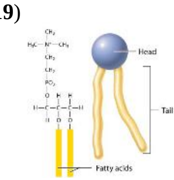


Table 2.3 Organic Molecules.

Lipids

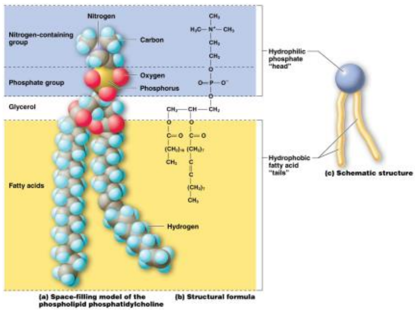
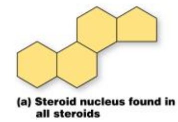


Figure 2.19 Lipids: structure of phospholipids.

Lipids

Steroids – nonpolar and share a *four-ring hydrocarbon structure* called the **steroid nucleus**



Cholesterol – steroid that forms basis for all other steroids

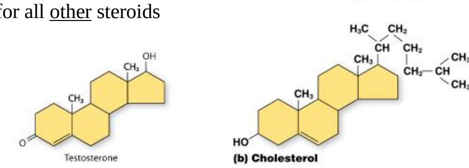


Figure 2.20 Lipids: structure of steroids and **Table 2.3** Organic Molecules.

Proteins

- **Proteins** are *macromolecules* that:
  - Function as *enzymes*
  - Play *structural* roles
  - Are involved in *movement*
  - Function in the body's *defenses*
  - Can be used as *fuel*

Proteins

- Twenty different **amino acids** (monomers of all proteins); can be linked by **peptide bonds** into **polypeptides**

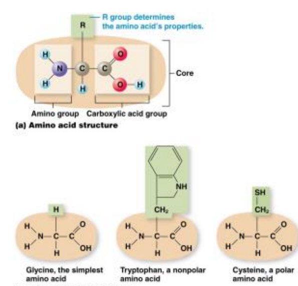


Figure 2.21a, b Proteins: structure of amino acids.

Proteins

Peptides – formed from two or more amino acids linked together by peptide bonds through **dehydration synthesis**:

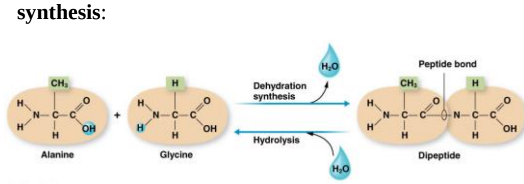


Figure 2.22 Proteins: formation and breakdown of dipeptides.

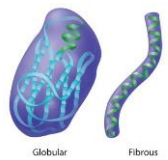
Proteins

- **Dipeptides** consist of *two* amino acids, **tripeptides** *three* amino acids, and **polypeptides** contain *10 or more* amino acids
- **Proteins** consist of *one or more* polypeptide chains folded into *distinct structures* which must be maintained to be functional; example of **Structure-Function Core Principle**

PROTEINS

Two basic types of proteins classified according to structure: **fibrous** and **globular**

- **Fibrous proteins** – long rope-like strands; composed mostly of *nonpolar* amino acids; link things together and add strength and durability to structures



- **Globular proteins** – spherical or globe-like; composed mostly of *polar* amino acids; function as enzymes, hormones, and other cell messengers

Table 2.3 Organic Molecules.

PROTEINS

Complex structure of a complete protein is divided into *four levels*:

- **Primary structure** – *amino acid sequence* of polypeptide chain

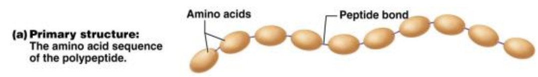


Figure 2.23a Levels of protein structure.

PROTEINS

Complex structure of a complete protein (continued):

- **Secondary structure** – one or more segments of primary structure *folded in specific ways*; held together by *hydrogen bonds*
  - **Alpha helix** – coiled spring
  - **Beta-pleated sheet** – Venetian blind

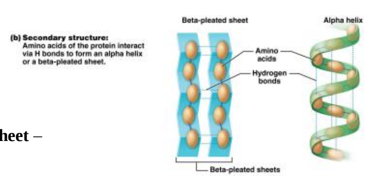


Figure 2.23b Levels of protein structure.

PROTEINS

Complex structure of a complete protein (continued):

- **Tertiary structure** – three-dimensional shape that peptide chain assumes (twists, folds, and coils including secondary structure); stabilized by *hydrogen bonding*

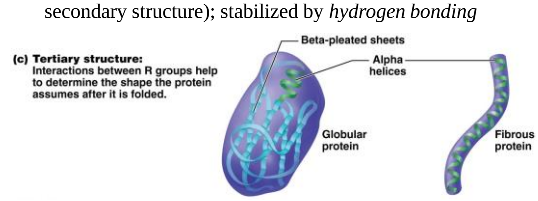


Figure 2.23c Levels of protein structure.

PROTEINS

Complex structure of a complete protein (continued):

- **Quaternary structure** – linking together *more than one polypeptide chain* in a specific arrangement; critical to function of protein as a whole

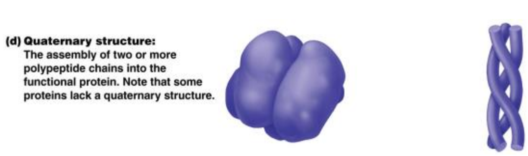


Figure 2.23d Levels of protein structure.

PROTEINS

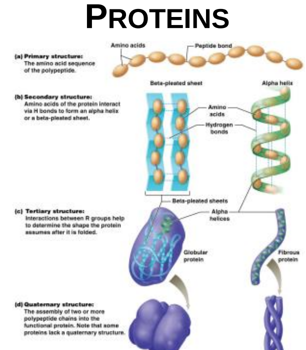


Figure 2.23 Levels of protein structure.

Proteins

- Protein **denaturation** – process of destroying a protein's *shape* by heat, pH changes, or exposure to chemicals
- <u>Disrupts</u> hydrogen bonding and ionic interactions that stabilize *structure* and *function.*

Nucleotides and Nucleic Acids

- **Nucleotides** – monomers of **nucleic acids**; named because of abundance in *nuclei* of cells; make up *genetic material*
- Nucleotide structure:
  - **Nitrogenous base** with a *hydrocarbon ring* structure
  - Five-carbon **pentose** sugar, **ribose** or **dexoyribose**
  - **Phosphate** group

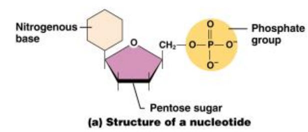


Figure 2.24a Structure of nucleotides.

Nucleotides and Nucleic Acids

Two types of nitrogenous bases: **purines** and **pyrimidines**

- <u>Purines</u> – *double-ringed* molecule; **adenine** (A) and **guanine** (G)
- <u>Pyrimidines</u> – single-ringed *molecule*; **cytosine** (C), **uracil** (U) and **thymine** (T)

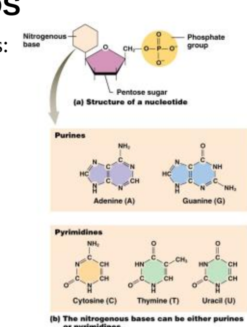


Figure 2.24 Structure of nucleotides.

Nucleotides and Nucleic Acids

Adenosine triphosphate (**ATP**)

- Adenine attached to ribose and <u>three</u> phosphate groups; main source of *chemical energy* in body
- Synthesized from **adenosine diphosphate** (**ADP**) and a **phosphate group** (**Pi**) using energy from oxidation of fuels (like glucose)

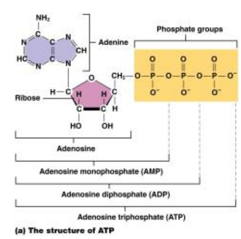


Figure 2.25a Nucleotides: structure and formation of ATP.

Nucleotides and Nucleic Acids

Adenosine triphosphate (continued):

- Potential energy in this "high-energy" bond can be released as **kinetic energy** to do **work**
- Production of large quantities of ATP <u>requires</u> oxygen; *why we breathe air*

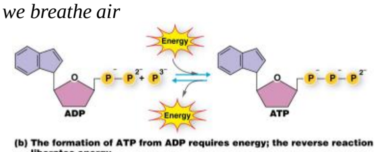


Figure 2.25b Nucleotides: structure and formation of ATP.

Nucleotides and Nucleic Acids

DNA, an extremely large molecule found in *nuclei* of cells; composed of <u>two</u> long chains that twist around each other to form a **double helix**

DNA contains **genes** – provide **recipe** or **code** for **protein synthesis** – process of making <u>every</u> protein

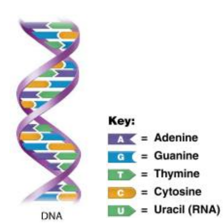


Figure 2.26a Structure of nucleic acids and Table 2.3 Organic Molecules.

Nucleotides and Nucleic Acids

Other *structural features* of DNA include:

- DNA contains:
  - Pentose sugar **deoxyribose** (lacks oxygen-containing group of ribose) forms *backbone of strand*; alternates with **phosphate group**
  - Bases: **adenine**, **guanine**, **cytosine**, and **thymine**

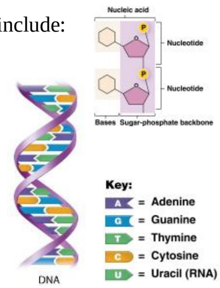


Figure 2.26a Structure of nucleic acids and **Table 2.3** Organic Molecules.

Nucleotides and Nucleic Acids

Other *structural features* of DNA include:

- Double helix strands – held together by hydrogen bonding *between the bases* of each strand
- Each base faces the *inside* of the double helix as strands run in <u>opposite</u> directions.

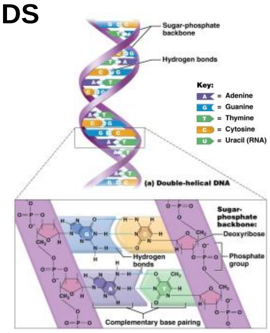


Figure 2.26a Structure of the nucleic acids DNA and RNA.

Nucleotides and Nucleic Acids

Other structural features of DNA (continued):

- DNA exhibits *complementary base pairing*; purine A <u>always</u> pairs with pyrimidine T and purine G <u>always</u> pairs with pyrimidine C
- **A = T** (where = denotes 2 hydrogen bonds) and **C ≡ G** (where ≡ denotes 3 hydrogen bonds)

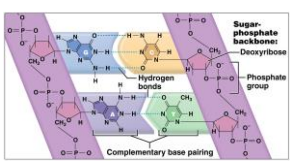


Figure 2.26a Structure of the nucleic acids DNA and RNA.

Nucleotides and Nucleic Acids

RNA – <u>single</u> strand of nucleotides; can move between nucleus of cell and cytosol; critical to making **proteins**

- RNA contains the *pentose sugar* **ribose**
- RNA contains **uracil** <u>instead</u> of **thymine**; still pairs with adenine, (A = U)

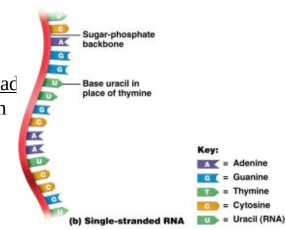


Figure 2.26b Structure of the nucleic acids DNA and RNA.

Nucleotides and Nucleic Acids

RNA –single strand of nucleotides (continued)

- RNA copies *recipe for specific protein* (**gene** in DNA); process called **transcription**
- RNA exits nucleus to protein synthesis location; then *directs the making of protein* from recipe; process called **translation**

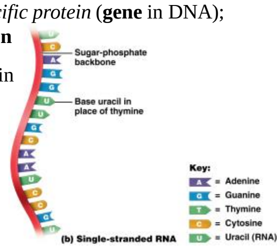


Figure 2.26b Structure of the nucleic acids DNA and RNA.

Nucleotides and Nucleic Acids

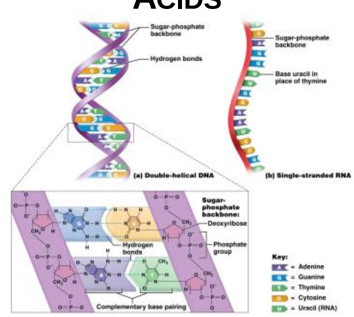


Figure 2.26 Structure of the nucleic acids DNA and RNA.