Physiology
Unit 1

CHEMISTRY REVIEW
Matter and Energy

• Definitions

• Types of energy
  – Kinetic vs. potential

• Forms of energy
  – Chemical
    • Ex: ATP
  – Electrical
    • Ex: Action potential of an neuron
  – Mechanical
    • Ex: Action of muscles in moving body
  – Radiant (Electrochemical)
    • Ex: Light energy
Atoms

• Smallest unit of matter
  – Composed of subatomic particles
    • Protons carry a positive charge
    • Neutrons are electrically neutral
    • Electrons carry a negative charge

• Nucleus is composed of protons and neutrons

• Electrons orbit the nucleus
  – Shells
    • Fixed distance from the nucleus
    • 2, 8, 18, 32
  – Orbitals
    • The direction the electrons move around the nucleus
    • can hold 2 electrons
    • Be able to write electron configurations
Elements = every specific type of atom
Atomic number: \# protons
Atomic mass: \#protons + \# neutrons
Chemical Composition of the Body

• 117 elements (as of 2006)
• 92 naturally occurring
• 24/112 are essential
  – 7 essential minerals
    • Ca, P, K, S, Na, Cl, Mg
  – 13 essential trace elements
    • Fe, I, Cu, Zn, Mn, Co, Cr, Se, Mo, F, Sn, Si, V
• 99% of the body’s atoms
  – Carbon - Oxygen
  – Hydrogen - Nitrogen
Ions

• Atom that has gained or lost one or more valence electrons

• electrolytes

• anions vs. cations
  – Na\(^+\) (atomic number 11)
  – Cl\(^-\) (atomic number 17)
  – H\(^+\) (atomic number 1)

• Minerals
  – Mg\(^{2+}\), Ca\(^{2+}\), Fe\(^{2+}\), SO\(_4\)\(^{2-}\)

• Trace elements
Isotopes

• Elements with a different number of neutrons
• Radioisotopes
  – Instability and disintegration of atomic nucleus
  – Occurs in heavier isotopes
  – Half-life = amount of time required for ½ of radioactivity to be lost as isotope disintegrates
  – Applications
    • Radioactive tracers
    • Radiation therapy for cancer
Valence Electrons

• Shells surround electrons
  – First shell can contain only 2 electrons.
    • If more than 2 electrons, must occupy shells more distant
  – Second shell can contain 8 electrons

• Valence electrons
  – Electrons in the outer most orbital that participate in chemical reactions (if orbit incomplete)
    • Form chemical bonds
Free Radicals

- Atoms containing a single unpaired valance electron
- Highly reactive
- Formed by specific enzymes, UV radiation, smoking, pollution
- **The Good**...Begins process of pathogen destruction by white blood cells
- **The Bad**...Can damage self DNA or cell membranes
- To prevent free radical damage the body has a defense system of *antioxidants*
  - Vitamins E, C, A (derived from beta carotene)
Chemical Bonds

• Chemical bonds:
  – Interaction of valence electrons between 2 or more atoms.

• Number of bonds determined by number of electrons needed to complete outermost shell.
Chemical bonds link atoms together to produce molecules

- Covalent
- Polar covalent
- Ionic
- Hydrogen
- van der Waals forces
Covalent bonds

• Atoms share electrons
• Most prevalent in organic compounds
• Non-polar = no charge

<table>
<thead>
<tr>
<th></th>
<th>Neutrons</th>
<th>Protons</th>
<th>Electrons</th>
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<tbody>
<tr>
<td>Carbon</td>
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<td>6</td>
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<tr>
<td>Hydrogen</td>
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</table>

Methane (four covalent bonds)
Polar covalent bonds

H₂O

- Unequal sharing of electrons
- Result = partial positive and partial negative poles
- The box shows a molecule of H₂O
- Polar covalent bonds make hydrogen bonds possible
  - Oxygen, nitrogen, phosphorous have tendency to pull electrons towards themselves.
Ionic bonds

- Complete transfer of one or more electrons from one atom to another
- Form ions when dissociate

\[ \text{Na}^+ \quad \text{Cl}^- \]
Hydrogen Bonds

Water

- Electrical attraction between hydrogen in one polarized bond and oxygen or nitrogen on another or in the same molecule
- Surface tension, cohesion
- Hydrogen bonding forms water
Van der Waals Forces

• Very weak attractions between nonpolar regions of molecules
• Important in protein structure and structure of other large molecules
• Caused by hydrophobic interactions
  – polar groups turn outward toward aqueous solution
  – Non-polar groups turn inward towards each other
Solutions

- Solutes and solvents
- Solubility
  - Hydrophilic
  - Hydrophobic
  - Amphipathic

\[
\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{L solution}}
\]
Solubility

- Glucose, amino acids, are H$_2$O soluble.
  - Hydration spheres form around atoms of oxygen, nitrogen, phosphorous
  - Charged complex ions and their cations form hydration spheres
- **Hydrophilic molecules**

- Molecules composed of non-polar covalent bonds are not H$_2$O soluble
  - Lipid soluble
  - Cannot form hydration spheres
  - **Hydrophobic molecules**
Acids, Bases, and pH

- Acids = proton donors
- Bases = proton acceptors
- pH = -log [ H⁺ ]
  - pH scale
    - 0 (most acidic) to 14 (most basic)
    - 0-6.9 is acidic
    - 7 is neutral
    - 7.1-14 is basic
  - more free H⁺ in solution = lower pH = more acidic!!!
### Table 2.3 The pH Scale

<table>
<thead>
<tr>
<th>H⁺ Concentration (Molar)*</th>
<th>pH</th>
<th>OH⁻ Concentration (Molar)*</th>
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*Molar concentration is the number of moles of a solute dissolved in one liter. One mole is the atomic or molecular weight of the solute in grams. Since hydrogen has an atomic weight of one, one molar hydrogen is one gram of hydrogen per liter of solution.
Buffers
Bicarbonate Buffer System

• System of molecules and ions that act to prevent changes in \([H^+]\).
• Stabilizes pH of a solution.
• In blood:
  • \(H_2O + CO_2 \rightleftharpoons H_2CO_3 \rightleftharpoons H^+ + HCO_3^-\)
    – Reaction can proceed in either direction (depending upon the concentration of molecules and ions).
    – This is the bicarbonate reaction
    – Be able to write this out