Summary: Chapter 2-Chemistry Review

I. Bonding

A. How many bonds between C, O, N, and H

B. Geometry
   1. Tetrahedral
   2. Planar - e.g. C forms double bond with C or O

C. Electronegativity
   1. (N & O) > (C & H)
   2. Bonds between C or H and O or N are polar; bonds between C & H are not

D. H₂O
   1. is more electronegative ➔ Polar Bonds
   2. Hydrogen Bonds attract molecules
   3. dissociation of H₂O = H⁺ + OH⁻

E. Ionic bonds -- e.g. Na⁺ and Cl⁻

F. van der Waals Interactions
Elements & Compounds

- An **Element** is a substance that cannot be broken down to other substances by chemical means.

- A **Compound** is a substance consisting of two or more elements in **fixed ratios**.

- A compound has characteristics different from those of its elements.
Elements & Compounds

Atoms (Elements)

Molecules (Compounds)

Cells
**Fig. 2.5: Simplified model of a Helium (He) Atom**

Cloud of negative charge (2 electrons)

Nucleus

Electrons

Atomic number

= number of protons

2

He

Helium

4.002602

Mass Number (~atomic mass)

= number of Neutrons + Protons

= 4 for Helium

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The Periodic Table of the Elements

These Elements make up ~97% of living organisms

For elements with no stable isotopes, the mass number of the isotope with the longest half-life is in parentheses.

Atoms of different elements differ in their number of subatomic particles:

<table>
<thead>
<tr>
<th>Element</th>
<th>Protons</th>
<th>Neutrons</th>
<th>Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>1</td>
<td>0</td>
<td>H</td>
</tr>
<tr>
<td>C</td>
<td>6</td>
<td>6</td>
<td>C</td>
</tr>
<tr>
<td>N</td>
<td>7</td>
<td>7</td>
<td>N</td>
</tr>
<tr>
<td>O</td>
<td>8</td>
<td>8</td>
<td>O</td>
</tr>
<tr>
<td>Na</td>
<td>11</td>
<td>23</td>
<td>Na</td>
</tr>
<tr>
<td>P</td>
<td>15</td>
<td>31</td>
<td>P</td>
</tr>
</tbody>
</table>

**Isotopes** of an element have the same number of protons but different number of neutrons:

- Stable isotopes:
  - $^{12}_{6}$C
  - $^{13}_{6}$C
- Unstable (radioactive): $^{14}_{6}$C
Chemical behavior of an atom depends mostly on number of electrons in **outermost shell** called **Valence Electrons**
Electrons in an atom vary in the amount of **potential energy** they possess (fixed, discrete amounts)

Different discrete energy levels correlate with average distance of electron from nucleus (electron shells)

(a) A ball bouncing down a flight of stairs provides an analogy for energy levels of electrons
We Are Carbon-Based Lifeforms

~96 % of the Human Body is:
**Oxygen (O)**
**Carbon (C)**
**Hydrogen (H)**
**Nitrogen (N)**

~4 %: Ca, P, K, S, Na, Cl, Mg

Trace elements (< 0.01%): B, Cr, Co, Cu, F, I, Fe, Mn, Mo, Se, Si, Sn, V, Zn

Composition of other organisms on Earth is similar
The first shell has only an **s orbital** that is **spherical** in shape.

The second shell has another **s orbital** and **3 p orbitals** that are each shaped somewhat like dumbbells aligned along 3 orthogonal axes.

Each orbital can hold a maximum of **2 electrons**.

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**Neon, with two filled Shells (10 electrons)**

**(a) Electron distribution diagram**

<table>
<thead>
<tr>
<th>First shell</th>
<th>Second shell</th>
</tr>
</thead>
<tbody>
<tr>
<td>1s orbital</td>
<td>2s orbital</td>
</tr>
<tr>
<td></td>
<td>Three 2p orbitals</td>
</tr>
</tbody>
</table>

**(b) Separate electron orbitals**

1s, 2s, and 2p orbitals

**(c) Superimposed electron orbitals**
Fig. 2.11: Chemical Bonds Link Atoms Together

**Covalent bonds**

- Form when two atomic nuclei share one or more pairs of electrons filling their **orbitals**.
- The **orbitals** have orientations in space that give molecules three-dimensional shapes.
- In H₂ – Each H atom is able to fill its 1s orbital by sharing its single electron with the other H atom.
Bonding properties of the most common elements in biological molecules

**H**: Atomic No. 1 ➔ 1 electron: \(1s^1\)
1st shell needs 1 electron ➔ Forms 1 covalent bond

**C**: Atomic No. 6 ➔ 6 electrons: \(1s^2 2s^2 2p_x^1 2p_y^1 2p_z^0\)
2nd shell needs 4 electrons ➔ Forms 4 covalent bonds

**N**: Atomic No. 7 ➔ 7 electrons: \(1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1\)
2nd shell needs 3 electrons ➔ Forms 3 covalent bonds

**O**: Atomic No. 8 ➔ 8 electrons: \(1s^2 2s^2 2p_x^2 2p_y^1 2p_z^1\)
2nd shell needs 2 electrons ➔ Forms 2 covalent bonds
The orbitals used by C to form 4 bonds to 4 different atoms are \textit{sp}^3 hybrid atomic orbitals, a combination of the 2s, 2p_x, 2p_y, and 2p_z orbitals. \textit{sp}^3 orbitals point to the corners of a Tetrahedron.

This is called “\textit{Tetrahedral Geometry}” characterized by bond angles of 109°.
A covalent bond that has electron density from shared electrons between the atoms is called a **σ bond**

The atoms in water and methane are connected by **σ bonds**.

### Fig. 2.17b: Molecular Shape Models

<table>
<thead>
<tr>
<th></th>
<th>Space-Filling Model</th>
<th>Ball-and-Stick Model</th>
<th>Hybrid-Orbital Model (with ball-and-stick model superimposed)</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Water (H₂O)</strong></td>
<td><img src="image" alt="Water Model" /></td>
<td><img src="image" alt="Hydrogen Model" /></td>
<td><img src="image" alt="Hydrogen Model" /></td>
</tr>
<tr>
<td><strong>Methane (CH₄)</strong></td>
<td><img src="image" alt="Methane Model" /></td>
<td><img src="image" alt="Carbon Model" /></td>
<td><img src="image" alt="Carbon Model" /></td>
</tr>
</tbody>
</table>

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**(b) Molecular-shape models**

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Fig. 2.12: Covalent Bonding in Four Molecules

<table>
<thead>
<tr>
<th>Name and Molecular Formula</th>
<th>Electron Distribution Diagram</th>
<th>Lewis Dot Structure and Structural Formula</th>
<th>Space-Filling Model</th>
</tr>
</thead>
<tbody>
<tr>
<td>(a) Hydrogen (H₂)</td>
<td><img src="image" alt="Electron Distribution Diagram" /></td>
<td>H: H</td>
<td><img src="image" alt="Space-Filling Model" /></td>
</tr>
<tr>
<td></td>
<td></td>
<td>H—H</td>
<td></td>
</tr>
<tr>
<td>(b) Oxygen (O₂)</td>
<td><img src="image" alt="Electron Distribution Diagram" /></td>
<td>O::O</td>
<td><img src="image" alt="Space-Filling Model" /></td>
</tr>
<tr>
<td></td>
<td></td>
<td>O=O</td>
<td></td>
</tr>
<tr>
<td>(c) Water (H₂O)</td>
<td><img src="image" alt="Electron Distribution Diagram" /></td>
<td>:O:H</td>
<td><img src="image" alt="Space-Filling Model" /></td>
</tr>
<tr>
<td></td>
<td></td>
<td>O—H</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>H—H</td>
<td></td>
</tr>
<tr>
<td>(d) Methane (CH₄)</td>
<td><img src="image" alt="Electron Distribution Diagram" /></td>
<td>H: C: H</td>
<td><img src="image" alt="Space-Filling Model" /></td>
</tr>
<tr>
<td></td>
<td></td>
<td>H:H</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>H—C—H</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>H—H</td>
<td></td>
</tr>
</tbody>
</table>
The two C atoms in Ethene form a second or double bond between themselves using their third unhybridized 2p orbitals.

The second bond is called a $\pi$ bond and prevents the C atoms from rotating around the bonds connecting them.

Note: the $\pi$ bond does not place electron density between the atoms.
Electronegativity is a measure of how strongly an atom attracts electrons.

<table>
<thead>
<tr>
<th>Element</th>
<th>Symbol</th>
<th>Electronegativity</th>
</tr>
</thead>
<tbody>
<tr>
<td>Oxygen</td>
<td>O</td>
<td>3.5</td>
</tr>
<tr>
<td>Chlorine</td>
<td>Cl</td>
<td>3.1</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>N</td>
<td>3.0</td>
</tr>
<tr>
<td>Carbon</td>
<td>C</td>
<td>2.5</td>
</tr>
<tr>
<td>Phosphorous</td>
<td>P</td>
<td>2.1</td>
</tr>
<tr>
<td>Hydrogen</td>
<td>H</td>
<td>2.1</td>
</tr>
<tr>
<td>Sodium</td>
<td>Na</td>
<td>0.9</td>
</tr>
<tr>
<td>Potassium</td>
<td>K</td>
<td>0.9</td>
</tr>
</tbody>
</table>

- **Nonpolar** covalent bonds form when the **Electronegativities** of two atoms are approximately equal.
- **Polar** covalent bonds form between atoms with strong electronegativity (such as oxygen) bonded to atoms with weaker electronegativity (such as hydrogen).
- In a **polar covalent bond** one atom has a partial **positive** charge or $\delta^+$ and the other atom has a partial **negative** charge or $\delta^-$. 
Fig. 2.13: Polar Covalent Bonds Are Formed Between Atoms With Unequal Electronegativity

Partial negative charge on oxygen

Partial positive charge on each hydrogen
Fig. 2.17b: $sp^3$ Hybrid Orbitals in Water, $H_2O$

<table>
<thead>
<tr>
<th>Space-Filling Model</th>
<th>Ball-and-Stick Model</th>
<th>Hybrid-Orbital Model (with ball-and-stick model superimposed)</th>
</tr>
</thead>
</table>

Water ($H_2O$)

Unbonded Electron pair
Hydrogen Bonds are attractions between the partial negative charge on the Oxygen atoms and the partial positive charges on the Hydrogen Atoms.

This gives water special properties that we will discuss next time.
Fig. 2.16: **Hydrogen Bonds** can form between water and other polar molecules.
**Fig. 2.14: Ionic Bonds**

Electronegativity:  \( \text{Na} = 0.9 \) &  \( \text{Cl} = 3.1 \)

Electrically charged **ions** form when an atom gains or loses one or more electrons to an atom of **very** different **electronegativity**.

**Ionic Bonds** are electrical attractions between oppositely charged ions.
Solvation: Water molecules surround ions and neutralize their charge.

Sodium (Na)
Chlorine (Cl)
Weak Chemical Bonds: van der Waals Interactions

Weak attraction of atoms due to constant motion of electrons:
## Chemical Bonds and Interactions

<table>
<thead>
<tr>
<th>NAME</th>
<th>BASIS OF INTERACTION</th>
<th>STRUCTURE</th>
<th>BOND ENERGY(a) (KCAL/MOL)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Covalent bond</td>
<td>Sharing of electron pairs</td>
<td>![Covalent Bond]</td>
<td>50–110</td>
</tr>
<tr>
<td>Hydrogen bond</td>
<td>Sharing of H atom</td>
<td>![Hydrogen Bond]</td>
<td>3–7</td>
</tr>
<tr>
<td>Ionic bond</td>
<td>Attraction of opposite charges</td>
<td>![Ionic Bond]</td>
<td>3–7</td>
</tr>
<tr>
<td>Hydrophobic interaction</td>
<td>Interaction of nonpolar substances in the presence of polar substances</td>
<td>![Hydrophobic Interaction]</td>
<td>1–2</td>
</tr>
<tr>
<td>van der Waals interaction</td>
<td>Interaction of electrons of nonpolar substances</td>
<td>![van der Waals Interaction]</td>
<td>1</td>
</tr>
</tbody>
</table>

\(a\)Bond energy is the amount of energy needed to separate two bonded or interacting atoms under physiological conditions.

- **High Energy**
- **Low Energy**
- **Very Low Energy**