UNIT 2: Electrons and where they are (Orbitals)

- Electrons are too small to locate directly, so scientists say that electrons “can be found” in specific areas around the nucleus. These areas are called _______________.
- These orbitals are complex mathematical functions, but we only need to use a few important numbers to describe them.
  1. **Principal quantum number**: Symbol \( n \)
     Represents the size and energy of an atomic orbital. Also called “Main energy level”
     \( n = 1, 2, 3... \)
  2. **Sublevel**: Symbols \( s, p, d, f \)
     Each main energy level has a certain number of sublevels.
     \( n = 1 \) only has 1 sublevel, which is called the ______ sublevel.
     \( n = 2 \) has 2 sublevels, which are called the ____, ____ sublevels.
     \( n = 3 \) has 3 sublevels, which are called the ____, ____ ____ sublevels.
     \( n = 4 \) has 4 sublevels, which are called the ____, ____ ____ ____ sublevels.
     **Each sublevel has a certain number of orbitals.**
     Every s sublevel has _____ orbital.
     Every p sublevel has __________ orbitals.
     Every d sublevel has _________ orbitals.
     Every f sublevel has ____________ orbitals.

**Energy of orbitals**: The energies of the sublevels within one main energy level are different according to this rule: \( s < p < d < f \) (highest energy)

This diagram on the right shows the relative energies of all the different orbitals.

- Electrons like to occupy, or “fill” lowest energy orbitals first. Electrons will occupy the next lowest energy orbital if the lower energy orbitals are full.
- Because it is hard to try and memorize the diagram as it is shown above, a useful way of memorizing the order of increasing energy orbitals is by using this method shown below.
  (Note that the picture on the right is a method of memorizing orbital energies.)
Electron Configurations

Electron configuration is a representation of electrons in an atom. (i.e. Which orbitals the electrons are in). You MUST follow THREE RULES when writing electron configurations.

1. **Aufbau Principle:**
   - Electrons occupy the lowest energy orbital available first.

2. **Pauli Exclusion Principle:**
   - Only two electrons can occupy any orbital at a time.

3. **Hund’s rule:**
   - Electrons occupy all the unoccupied orbitals of the same energy subshell first, before pairing up and filling the orbitals.

Example: The electron configuration for Nitrogen is \(1s^22s^22p^3\)

The orbital diagram for Nitrogen is \[
\begin{array}{c}
1s \\
2s \\
2p
\end{array}
\]

**Practice:** write the electron configurations for Argon, Neon, Phosphorus, Sulfur, Oxygen, Scandium.

**Valence Electrons:** The “valence shell” is the term used to describe the outermost main energy level for that particular atom.

For example, Nitrogen’s electron configuration is \(1s^22s^22p^3\), so it’s highest main energy level is \(n = 2\), (the second energy level) So nitrogen has _______ valence electrons. (i.e. Five electrons in the second energy level)

Valence electrons are the electrons that are involved in bonding, so they are important.

- **Lewis Dot Diagrams:** Show just the valence electrons of an \(\text{\text{N}}\) atom.

**Periodic Table**

The periodic table can be grouped by some representative characteristics.

- **Metals, Nonmetals, Metalloids:** Label the diagram below.

**Properties of Metals:** ________, good conductors of ________, and ________, are malleable.

**Properties of Nonmetals:** Usually brittle, poor conductors of ________, and ________.

**Properties of Metalloids:**

Metalloids are semiconductors, and have some characteristics of metals, some characteristics of nonmetals.

**Periodic Table classification by valence electrons:** (diagram on the right above)
Periodic Trends

The properties of atoms change in a pattern that is called a periodic trend.

- **Atomic Radius**: The size of an atom is described by its radius. (This is called atomic radius)
  The atomic radius decreases across a period (from left to right)
  This is because the number of electrons between the valence shell and the nucleus is the same for all elements in the same period, but since there is an increasing positive charge in the nucleus as you move across the period, the valence shell electrons get pulled in closer to the nucleus, making the overall atom smaller.

  The atomic radius increases down a group
  This is because electrons in the valence shells are in higher energy levels as you move down a group, which means they are further from the nucleus, which means that the atom is larger.

- **Ionic Radius**: _________ across a period, ____________ down a group

- **Ionization Energy**: The energy needed to remove an electron (from an atom in the gas state)
  _________ across a period, ____________ down a group
  o High Ionization energy means it requires a lot of energy to remove an electron.
  Low Ionization energy means it is easy to remove an electron.

- **Electronegativity**: The ability of an atom to attract electrons to itself when in a bond.
  _________ across a period, ____________ down a group.

UNIT 3: Ionic Bonding

- Atoms can form ions because all atoms want full s and p subshells. A full s and p subshell (of the same energy level) is also called a “complete octet”, and is a very stable electron arrangement.
- Atoms therefore lose or gain their valence electrons to achieve the electron configuration of the nearest noble gas.
- Ionic bonding usually occurs between metals and non-metals.

**Metal atoms**: When metal atoms form ions, they tend to _______ electrons to form ___________ cations.
- All group 1 metals: Lose their ns\(^1\) valence electron. (one electron)
  Forms a +1 ion. Eg. Na\(^+\)
- All group 2 metals: Lose their ns\(^2\) valence electrons. (two electrons)
  Forms a +2 ion. Eg. Mg\(^{2+}\)
- All group 13 metals: Lose their ns\(^2\) np\(^1\). (three electrons)
  Forms a +3 ion. Eg. Al\(^{3+}\)
- Transition metals can form more than one type of cation (because they can lose two/three/four/etc. electrons) Transition metals always lose their (n+1)s\(^2\) electrons first.
  Eg. Scandium loses its 4s\(^2\) before losing its 3d\(^1\) electron.
  The charge of a transition metal can be called its “oxidation number.”

**Nonmetal atoms**: When non-metal atoms form ions, they tend to _______ electrons to form ___________ anions.
- Group 17 nonmetals: Gain one electron to form a _______ ion. eg. F\(^-\)
- Group 16 nonmetals: Gain two electrons to form a _______ ion. eg. O\(^{2-}\)
- Group 15 nonmetals: Gain three electrons to form a _______ ion. eg. N\(^{3+}\)

**Ionic Bonding**: When two ions meet, they are brought close together by their opposite electric charges, which is called an ionic bond.
• **Ionic Compounds:** Compounds that are formed by ionic bonding have positive and negative ions arranged in an organized arrangement, which is called a *crystal lattice*.
  - Brittle
  - High melting and boiling points (because of strong attractions between charges)
  - Bigger charge = stronger ionic bond.
  - Smaller ions = stronger ionic bond.
  - Can conduct electricity when the ions can move (ie. When dissolved in water or when melted “molten”)

• **Formula Unit:** The formula for the smallest repeating unit in an ionic compound. aka “formula”
  1. Cross the number of the charges onto the subscripts (number at the bottom) of the *other* element.
     
  2. The formula should always have the cation first, and the anion last.
  3. For polyatomic ions, keep the polyatomic unit as a unit (do *not* change polyatomic ion’s subscripts)

• **Naming:** 1. Name the metal ion first, with the metal’s full name.
   (For transition metals, write the transition metal ion’s *charge number* in Roman numerals in brackets after it’s full name. Eg. Fe\(^{2+}\) in FeCl\(_2\) becomes *Iron (II)* chloride)
   
  2. Then name the nonmetal ion, with the suffix –ide for single element nonmetals.
   (For polyatomic anions, just write the polyatomic unit’s name.)

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**Polyatomic ions:** Memorize name, ion formula, and charge.

<table>
<thead>
<tr>
<th>Ion</th>
<th>Name</th>
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</thead>
<tbody>
<tr>
<td>NH(_4^+)</td>
<td>Ammonium</td>
</tr>
<tr>
<td>NO(_2^-)</td>
<td>Nitrite</td>
</tr>
<tr>
<td>NO(_3^-)</td>
<td>Nitrate</td>
</tr>
<tr>
<td>OH(^-)</td>
<td>Hydroxide</td>
</tr>
<tr>
<td>CO(_3^{2-})</td>
<td>Carbonate</td>
</tr>
<tr>
<td>HCO(_3^-)</td>
<td>Hydrogen carbonate</td>
</tr>
<tr>
<td>SO(_3^{2-})</td>
<td>Sulfite</td>
</tr>
<tr>
<td>SO(_4^{2-})</td>
<td>Sulfate</td>
</tr>
<tr>
<td>PO(_4^{3-})</td>
<td>Phosphate</td>
</tr>
<tr>
<td>HPO(_4^{2-})</td>
<td>Hydrogen phosphate</td>
</tr>
<tr>
<td>H(_2)PO(_4^-)</td>
<td>Dihydrogen phosphate</td>
</tr>
<tr>
<td>CH(_3)COO(^-) or C(_2)H(_3)O(_2^-)</td>
<td>Acetate</td>
</tr>
</tbody>
</table>

**Practice Naming Ionic Compounds:**

1. NH\(_4\)Cl: ______________________
2. Ag\(_2\)CO\(_3\): ______________________
3. Fe(OH)\(_3\): ______________________
4. Calcium phosphate: ______________________
5. Lead (IV) chloride: ______________________
6. Strontium acetate: ______________________
7. Magnesium bromide: ______________________
8. BaS: ______________________
9. Lithium hydrogen phosphate: ______________________
10. Calcium hydrogen carbonate: ______________________
Covalent Bonding

Nonmetal atoms can share valence electrons to acquire the stable electron configuration of noble gases (ie. an “octet”, or 8 valence electrons).

When two atoms share valence electrons, the two atoms are said to be covalently bonded to each other.

Compounds that are made up of covalent bonding are called molecules.

Covalent bonding usually occurs between nonmetals and nonmetals.

- One shared pair of electrons is called a __________ bond. Eg. F₂
- Two shared pairs of electrons is called a __________ bond. Eg. O₂
- Three shared pairs of electrons is called a __________ bond. Eg. N₂

A covalent bond created by direct orbital overlap is called a **sigma bond**.

A covalent bond created by sideways orbital overlap is called a **pi bond**.

- All single bonds (ie. one shared pair) are __________ bonds.
- All double bonds (ie. two shared pairs) have one __________ bond and one __________ bond.
- All triple bonds (ie. three shared pairs) have one __________ bond and two __________ bonds.

Practice: How many sigma bonds and pi bonds are there in the following molecules?

1. F₂
2. NH₃
3. H₂C≡C≡N

Naming Covalent compounds/Molecules:

- **Binary Molecular Compounds:**
  1. First element in the formula is named first, with its entire name.
  2. Second element has the suffix –ide.
  3. All elements have prefixes to indicate the number of that atom in the molecule.
  4. The first element never gets the “mono” prefix.
  5. Double vowels: the last vowel of the prefix is dropped.

- **Acids:**
  HF, HCl, HBr, HI: hydro __________ ic acid
  HNO₃: __________ acid
  H₂SO₄: __________ acid
  H₂S: __________ acid
  H₂CO₃: __________ acid

<table>
<thead>
<tr>
<th>Number of Atoms</th>
<th>Prefix</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>mono</td>
</tr>
<tr>
<td>2</td>
<td>di</td>
</tr>
<tr>
<td>3</td>
<td>tri</td>
</tr>
<tr>
<td>4</td>
<td>tetra</td>
</tr>
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<td>5</td>
<td>penta</td>
</tr>
<tr>
<td>6</td>
<td>hexa</td>
</tr>
<tr>
<td>7</td>
<td>hepta</td>
</tr>
<tr>
<td>8</td>
<td>octa</td>
</tr>
<tr>
<td>9</td>
<td>nona</td>
</tr>
<tr>
<td>10</td>
<td>deca</td>
</tr>
</tbody>
</table>

Write the formula of:

1. Dinitrogen pentoxide: ____________________
2. Carbon dioxide: ____________________
3. Hydrobromic acid: ____________________
4. Carbonic acid: ____________________
5. Carbon sulfide: ____________________

Write the name of:

6. SO₂: ____________________
7. N₂O: ____________________
8. NO₂: ____________________
9. SCl₆: ____________________
10. OF₂: ____________________
**Structural Formulas/Lewis Structures of Covalent Molecules**

1. The central atom is usually the one more to the LEFT of the periodic table. (Or it is the one with the smallest number in the formula)
2. Write down the sum of all the valence electrons available to bond.
3. Place one bonding pair between the central atom and each of the terminal atoms.
4. Subtract the number of bonding electrons from the total number of available electrons.
5. Fill the octets for all the terminal atoms first.
6. Then place any remaining electrons on the central atom.
7. If there aren’t enough electrons for the central atom to have an octet, then go back and convert one of the lone pairs on the terminal atoms into a double bond. Then check that everything still has an octet.
   - C, N, O, and S often form double or triple bonds.

**Polyatomic ions:** Eg. CO$_3^{2-}$ The atoms within a polyatomic ion are covalently bonded.
- If – charge: add electrons to the total number of valence electrons and then follow steps 1-7.
- If + charge: subtract electrons from the total number of valence electrons and then follow steps 1-7.

**Resonance structures:** Some molecules can have equivalent structures, eg. NO$_3^-$

**Exceptions to the Octet Rule:**

1. Some molecules have an **odd number of total valence electrons**.
   - NO$_2$, ClO$_2$, NO (*you will not be asked these in the exam)
2. Some atoms don’t need a full **octet** when drawing the Lewis structure.
   - This is **mainly Boron** molecules, and some Be and Li molecules (But Be and Li tend to form ionic compounds, not covalent compounds)
   - Eg. BH$_3$, BF$_3$
3. Some atoms can have an **expanded octet**, which means that they can have more than 8 electrons around themselves when in a covalent compound.
   - These atoms are usually the nonmetals in Period 3 and below. (eg phosphorous, P, sulfur S, etc.) Eg. PCl$_5$, SF$_6$

**Molecular Shapes:** **VSEPR theory:**
1. Draw Lewis structure.
2. Count the number of electron domains around the central atom.
3. Determine Electron domain geometry and then molecular geometry (Memorize table on this page)

To explain molecular shapes, we use the theory of **hybridization**.
- If there are 2 domains around the central atom, the hybridization is **sp**.
- If there are 3 domains around the central atom, the hybridization is **sp$^2$**.
- If there are 4 domains around the central atom, the hybridization is **sp$^3$**.
- If there are 5 domains around the central atom, the hybridization is **sp$^3$d**.
- If there are 6 domains around the central atom, the hybridization is **sp$^3$d$^2$**.

Practice:
What is the molecular geometry and hybridization of the central atom of:
1. NH$_3$
2. SO$_2$
3. PO$_4^{3-}$
4. CO$_2$
5. H$_2$O
6. PCl$_3$
7. PCl$_5$
8. N$_2$O
9. CCl$_4$
10. SF$_6$