Valence electrons are the electrons in the outer shell of an atom. They are important because they indicate the atom’s reactivity with other atoms.

The outermost electrons have the most energy (remember the ladder!)

**Valence electrons is the magic NUMBER.** Atoms are most stable when they have 8 electrons in their outer shell.

**Then both Li and F have FULL valence shells.**

**LEWIS DOT DIAGRAMS**

A Lewis Dot Diagram shows only the valence electrons of an atom.

For elements 1 through 20, there is a secret trick for knowing how many valence electrons an atom has. Look at the element’s group NUMBER on the Periodic Table.

Group 1 elements have 1 valence electron, group 2 elements have 2 valence electrons, and for groups 13 through 18:

- **Drop the one** from the number to determine the number of valence electrons.
- For example, carbon is in group 14 so it has 4 valence electrons.

**Li**

**F**
Homework

Assignment #5: Lewis Diagrams Worksheet pg. 26 & Practice Questions pg. 37

Lewis Diagrams / Bonding Worksheet

1. Draw electron dot diagrams for the following atoms (1 mark each)
   a. C
   b. H
   c. Na
   d. Cl
   e. P
   f. K
   g. Fr
   h. Mg
   i. Al
   j. O

2. Draw electron dot diagrams for the following ions (1 mark each)
   a. Na⁺
   b. Mg²⁺
   c. Si⁴⁺

3. Complete the following table by drawing both the Bohr diagram and the Lewis diagram for each element (assumes it is an atom and not an ion of that element).

<table>
<thead>
<tr>
<th>Name of Element</th>
<th>Bohr Diagram</th>
<th>Lewis Diagram</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbon</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Oxygen</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Lithium</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Chlorine</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Practice

Let's Try!

Write the group number above each group. Then draw the Lewis Dot Diagram for each element.

Why doesn't Helium follow the 'drop the one' rule?

1. This is a __________Diagram of a __________atom.
2. How many shells does this atom have? ______
3. What is the shell closest to the nucleus called? ______
4. How many electrons does this shell have? ______
5. Is this shell full? ______
6. What is the name of the second shell from the nucleus? ______
7. How many electrons does this shell have? ______
8. Is this shell full? ______
9. What is the name of the third shell from the nucleus? ______
10. How many electrons does this shell have? ______
11. Is this shell full? ______
12. Altogether, how many electrons does this atom have? ______
13. Is another shell needed? ______
14. How many more electrons would this atom need to have a full outer shell? ______

15. Draw the Lewis Dot Diagram for the atom: ______
16. How can you quickly double-check your diagram? ______
17. How many shells do 2 electrons need? ______
18. How many shells do 10 electrons need? ______
19. How many shells do 20 electrons need? ______

Write True or False to answer each statement:
20. Every atom has at least two electron shells: __________
21. The L shell always has 8 electrons: __________
22. If there is an L shell, it means that the K shell is full: __________

Have you feeling about the basics of Modeling Atoms? Create one Lewis electron: ______

Sum It Up!

15. Draw the Lewis Dot Diagram for the atom: O
16. How can you quickly double-check your diagram? ______
17. How many shells do 2 electrons need? ______
18. How many shells do 10 electrons need? ______
19. How many shells do 20 electrons need? ______

Write True or False to answer each statement:
20. Every atom has at least two electron shells: __________
21. The L shell always has 8 electrons: __________
22. If there is an L shell, it means that the K shell is full: __________

Have you feeling about the basics of Modeling Atoms? Create one Lewis electron: ______

Unit 2 Chemistry Page 16
Patterns of Electron Arrangement in Periods

The principal number of an element equals the number of occupied _shells_ of its atoms. For example, the first shell of a period 2 hydrogen atom contains 2 electrons, which are called the _nuclear shell_.

The first shell of an atom can have a maximum of 2 electrons.

The elements in period 2 all have two electron shells. For each element in period 2, the first shell, which is closest to the nucleus, has a _closed shell_. Another way to think about this is that the element has a _nuclear shell_. As you move from left to right across period 2, more electrons are added to the second shell of each atom. Notice that in the last element in period 2, each atom has a _fully occupied_ shell of its second shell.

Patterns of Electron Arrangement in Groups

The _valence shell_ is the highest occupied shell of an atom.

The electrons in the valence shell are called the _valence electrons_.

Valence electrons are involved in _chemical bonds_.

**Example:**
- The atoms of each element in group 1, hydrogen, have only _1_ electron in their valence shell.
- Group 2 elements have _2_ electrons in their valence shell.
- Group 17 elements have _7_ electrons in their valence shell.
- Group 18 elements have _8_ electrons, the noble gases, filling their valence shell.

**Note:** that elements can exist singly or in pairs. Electrons in _compulsively_ shell (except for hydrogen) appear in pairs.

FORMING COMPOUNDS

When two atoms come together, they bond to each other. A _chemical bond_ is formed between the atoms if the two arrangements of atoms and electrons are stable. The stability of an atom, ion, or compound is related to its _energy_. Compounds are more stable.

The _bond energy_ is achieved when the atoms in the compound have the same arrangement of valence electrons as a noble gas. A noble gas is an element that has the closest possible to one of three areas:
1. Atoms that gain _e⁻_ to other atoms, forming a _cation_.
2. Atoms that share _e⁻_ (ions from another atom, forming a _cation_).
3. Atoms that share _e⁻_ (ions from another atom, forming a _cation_).

**Types of Bonds**

A) **IONIC BOND:** An ionic compound contains a _positive ion_ (cation) and a _negative ion_ (anion) in the compound. One or more _electrons_ transfer from each atom of the metal to each atom of the non-metal.

**Lewis Structure:**

<table>
<thead>
<tr>
<th>Cation</th>
<th>Anion</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na⁺</td>
<td>Cl⁻</td>
</tr>
</tbody>
</table>

**Sodium Chloride:**

**Chemical Formula:**

**MgCl₂**

**Mg**₂⁺ + 2**Cl**⁻ → MgCl₂

**Ratio of Atoms:**

1 * **Mg**₂⁺ + 2 * **Cl**⁻ → MgCl₂

**Chemical Equation:**

**MgCl₂**
**Practice**

**EXAMPLES:**

1. sodium chloride: $NaCl$
2. potassium oxide: $K_2O$
3. calcium phosphate: $Ca_3(PO_4)_2$
4. tin(IV) sulphate: $Sn(IV)\,SO_4$

**CHALLENGE Problem** — Determining the Name of an Ionic Compound from Its Formula

**What to Think about**

1. Write the names of the two constituent ions.
2. Write the formulas of the possible compounds to see which one has the correct formula.

**How to Do It**

**Assignment #6:**

Write the formulas of the compounds contain the following ions:

1. Na$^+$ and Br$^-$
2. Zn$^{2+}$ and I$^-$
3. K$^+$ and $S^{2-}$
4. Al$^{3+}$ and O$^{2-}$
5. Ca$^{2+}$ and O$^{2-}$
6. Al$^{3+}$ and P$^{3-}$

Write the formulas of the following ionic compounds:

1. strontium nitride
2. lithium oxide
3. aluminium nitride
4. magnesium chloride

**B: COVALENT BOND:**

The atoms of every ___________ share electrons with other non-metal atoms.

In covalent bonding, atoms overlap slightly, and one n___________ electron from each atom will result in a pair of electrons sharing a ___________ bond.

Both atoms are attracted to the same pair of electrons, forming a ___________ bond.

A covalent ___________ is formed when two metallic atoms share electrons to form a covalent bond.

A covalent ___________ is a group of atoms in which the atoms are bound together by sharing one or more pairs of electrons.

The pair of electrons involved in a covalent bond are sometimes called the n___________ pair.

A pair of electrons in the n___________ shell that is not used in bonding is sometimes called a n___________ pair (optional).

**Bohr Diagram:**

![Bohr Diagram](image)

**Lewis Structure:**

**Properties of Ionic Compounds**

<table>
<thead>
<tr>
<th>Property</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>Electrical Conductivity</td>
<td>Good in solution, poor in solid</td>
</tr>
<tr>
<td>Thermal Conductivity</td>
<td>Poor</td>
</tr>
<tr>
<td>Color</td>
<td>Often bright or metallic</td>
</tr>
</tbody>
</table>

**Properties of Covalent Compounds**

<table>
<thead>
<tr>
<th>Property</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>Electrical Conductivity</td>
<td>Poor</td>
</tr>
<tr>
<td>Thermal Conductivity</td>
<td>Good in solid, poor in solution</td>
</tr>
<tr>
<td>Color</td>
<td>Often colorless, some have color</td>
</tr>
</tbody>
</table>

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Unit 2 Chemistry Page 18
DIATOMIC MOLECULES
You can use Lewis diagrams to help explain why some of the elements exist as diatomic molecules.
A ___________ molecule is a pair of atoms that are joined by ___________ bonds.
Diatomic elements form this way because the two atom molecules are _________ than the individual atoms.
For example, F₂ may be gas diatomic:
By joining together to form ________, each fluorine atom can achieve a full __________ shell of ________ electrons
Other diatomic elements are hydrogen (H₂), nitrogen (N₂), oxygen (O₂), chlorine (Cl₂), bromine (Br₂), and iodine (I₂).

The NOWBrICl Elements:
• There are 7 naturally occurring elements that, when found in nature, exist as ___________ molecules.
• To be stable, the atoms will not exist alone, they bond to each other.
• This means they must ________ be written with a subscript of ________

Assignment #7: “Check your Understanding.”
Questions #1-14, textbook pg 183