Small-Group Guided Inquiry:

**Lewis Structures for Covalent Molecules**

**Read This First!**

When two nonmetal atoms share electrons, they form covalent bonds. When a bond is formed, **each atom shares one of its electrons** with the other. So, **each bond is made of 2 shared electrons**. Two shared electrons are called a **single bond**; the remaining electrons can be called any of these terms: **lone pairs**, **nonbonding pairs**, **unpaired electrons**,

or **unshared electrons**.

**Part 1**

**Step 1:** *One molecule at a time*, use the puzzle pieces to build the correct Lewis structure for each covalent compound below. (*Hint!* You know you have the correct puzzle when all of the outside edges are straight.)

**Step 2:** From your completed puzzle, draw the molecule’s Lewis structure by replacing shared pairs of electrons with a bond (a short, straight line).

|  |  |  |  |
| --- | --- | --- | --- |
| **CH4** | **NH3** | **H2O** | **H2** |
| **PH3** | **CH2ClBr** | **H3CCH3** | **H2NNH2** |

Now, use what you have found from the puzzles and structures above to complete the chart below:

|  |  |  |  |
| --- | --- | --- | --- |
| **Element** | **How many bonds does this element typically form?** | **How many valence electrons does this element need to share in order to obtain an octet?** | **What column of the Periodic Table is this element located in?** |
| **Carbon** |  |  |  |
| **Nitrogen** |  |  |  |
| **Phosphorus** |  |  |  |
| **Oxygen** |  |  |  |
| **Hydrogen** |  |  |  |
| **Chlorine** |  |  |  |
| **Bromine** |  |  |  |

What patterns can you find in the chart above? Answer the following questions:

1. Use complete sentences to summarize what you can interpret from the chart that you just completed.
2. H, F, Cl, and Br are most likely to be terminal (ending) atoms rather than bridging (central) atoms in molecules. Why do we see this pattern?
3. If hydrogen and the halogens make good terminal atoms, what is one element that would make a good bridging atom? Explain your answer.

**Read This Next!**

Covalent bonds are formed when nonmetal atoms share electrons with each other. Atoms share electrons in order to obtain a full octet. When an atom gains its full valence shell, it becomes much more stable.

Often, it takes more than one electron to fill an atom’s valence shell. So, the atom will need **more than one bond** in order to be stable. Each bond is made of 2 shared electrons. Two shared electrons are called a single bond.

If atoms share more than 2 electrons, they form more bonds. **Atoms always share electrons in sets of 2** (one from each atom). Four shared electrons form **2 bonds**, called a **double bond**. Six shared electrons form **3 bonds**, called a **triple bond**. Usually, one pair of atoms never share more than 6 electrons. So, a triple bond is the **strongest bond** atoms can make.

Each pair of electrons is held together by intramolecular energy. **As the number of bonds increases, the amount of energy in the molecule increases.** As the number of bonds between any 2 atoms increases, the amount of energy they are sharing increases. Usually, this makes their bonds **stronger**. So, **a double bond is stronger than a single bond**, and **a triple bond is stronger than a double bond**.

Since most atoms are trying to obtain an octet (8 valence electrons), they can only form **up to 4 bonds**. Usually, **each individual atom can only have 4 total bonds**. If an atom has 4 bonds, it should not have any lone pairs of electrons; if an atom has less than 4 bonds, it should have lone pairs around it. The main **exceptions** are hydrogen (which can only have 2 total electrons) and boron (which can only have 6 total electrons).

**Part 2**

**Step 1:** *Just like earlier*, build each molecule using the puzzle pieces.

**Step 2:** Draw the correct Lewis structure for each molecule by turning each pair (2) of shared electrons into a bond. (*Remember!* You can only have up to 3 bonds – a triple bond – between any 2 atoms and each atom can only have up to 4 total bonds around it.)

|  |  |  |
| --- | --- | --- |
| **N2** | **HCN** | **H2CCH2** |
| **HClCCHCl** | **HNNH** | **HCOOH** |

Do the same patterns still apply? Answer the following questions:

1. If a double bond counts as 2 bonds and a triple bond counts as 3 bonds, do these molecules follow the same patterns you stated in Part 1, Question 1?
2. Explain the differences you see between the structures of each set of molecules below:

**Elemental hydrogen** (H2) and **elemental nitrogen** (N2)

**Ethane** (C2H6) and **Ethene** (C2H4)

**Diazane** (N2H4) and **Diazene** (N2H2)

**Bromochloromethane** (CH2ClBr), **Dichloroethene** (C2H2Cl2), and **Formic Acid** (CH2O2)

1. Name 2 patterns you can see in the **formal chemical names** given for some of the molecules above.
2. What is purpose of writing the chemical formula of a compound in different ways? (*For example*: **Ethane** can be written as **C2H6** or as **H3CCH3**.)