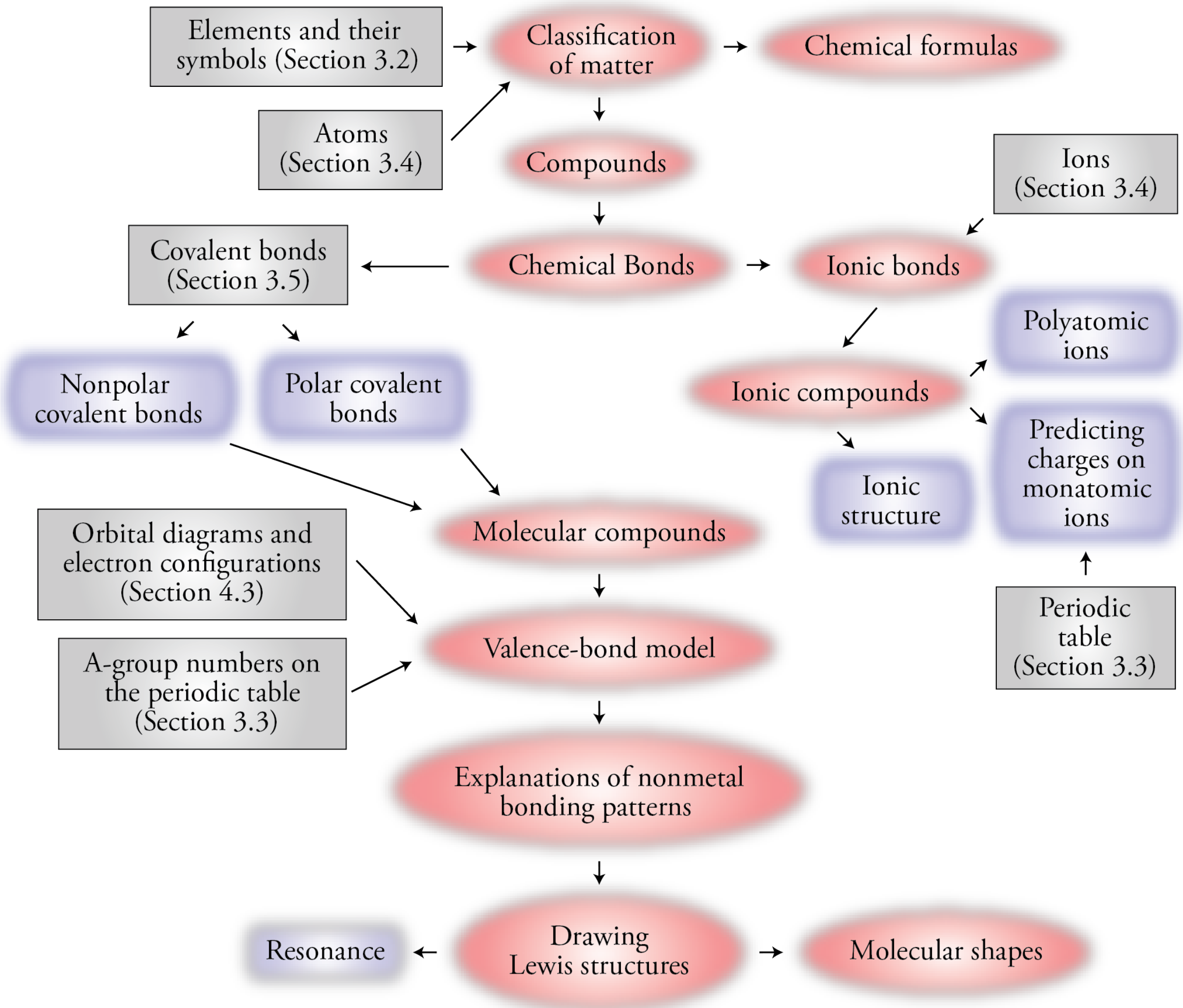
Chapter 5
Chemical Compounds

An Introduction to Chemistry

by Mark Bishop

Chapter Map



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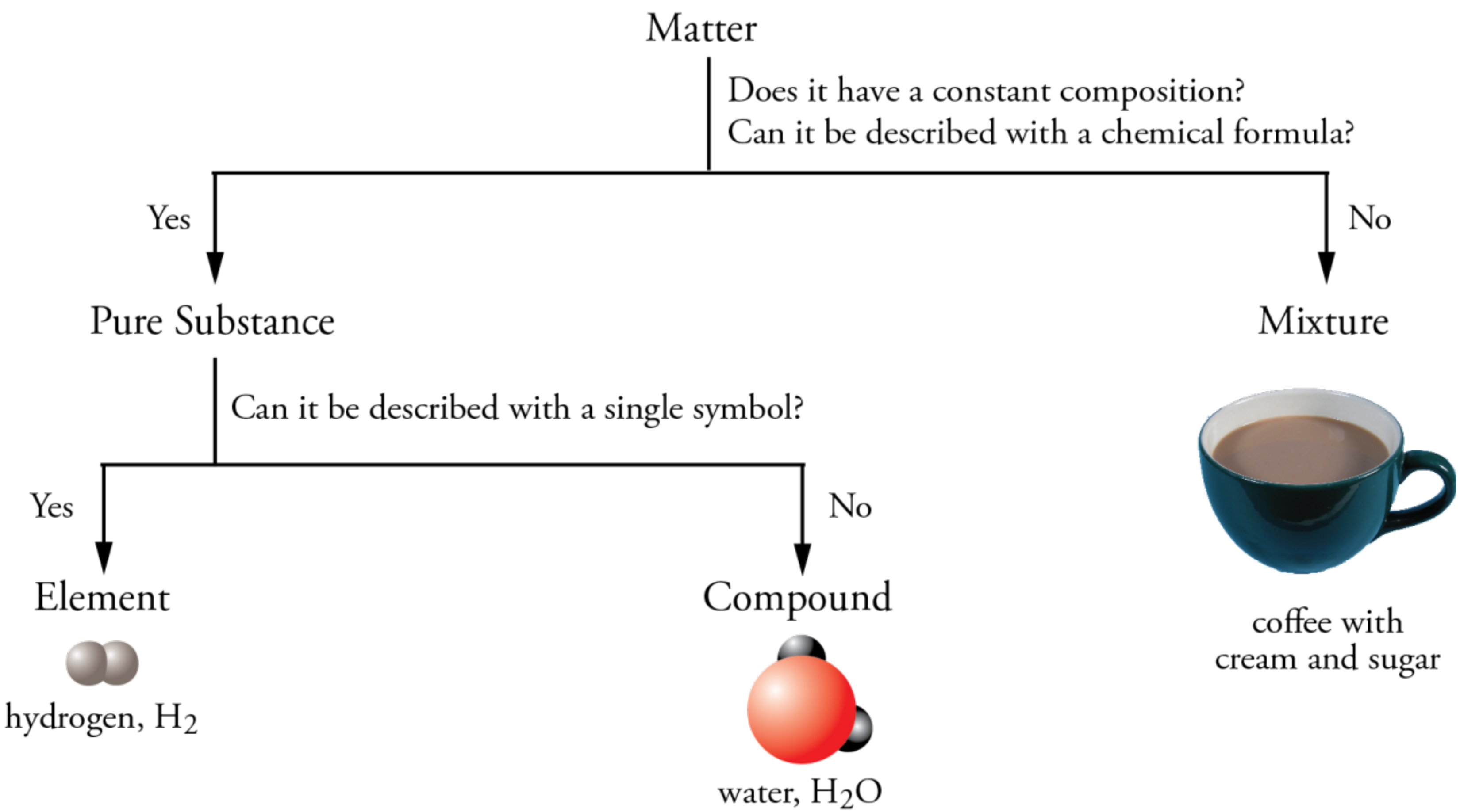
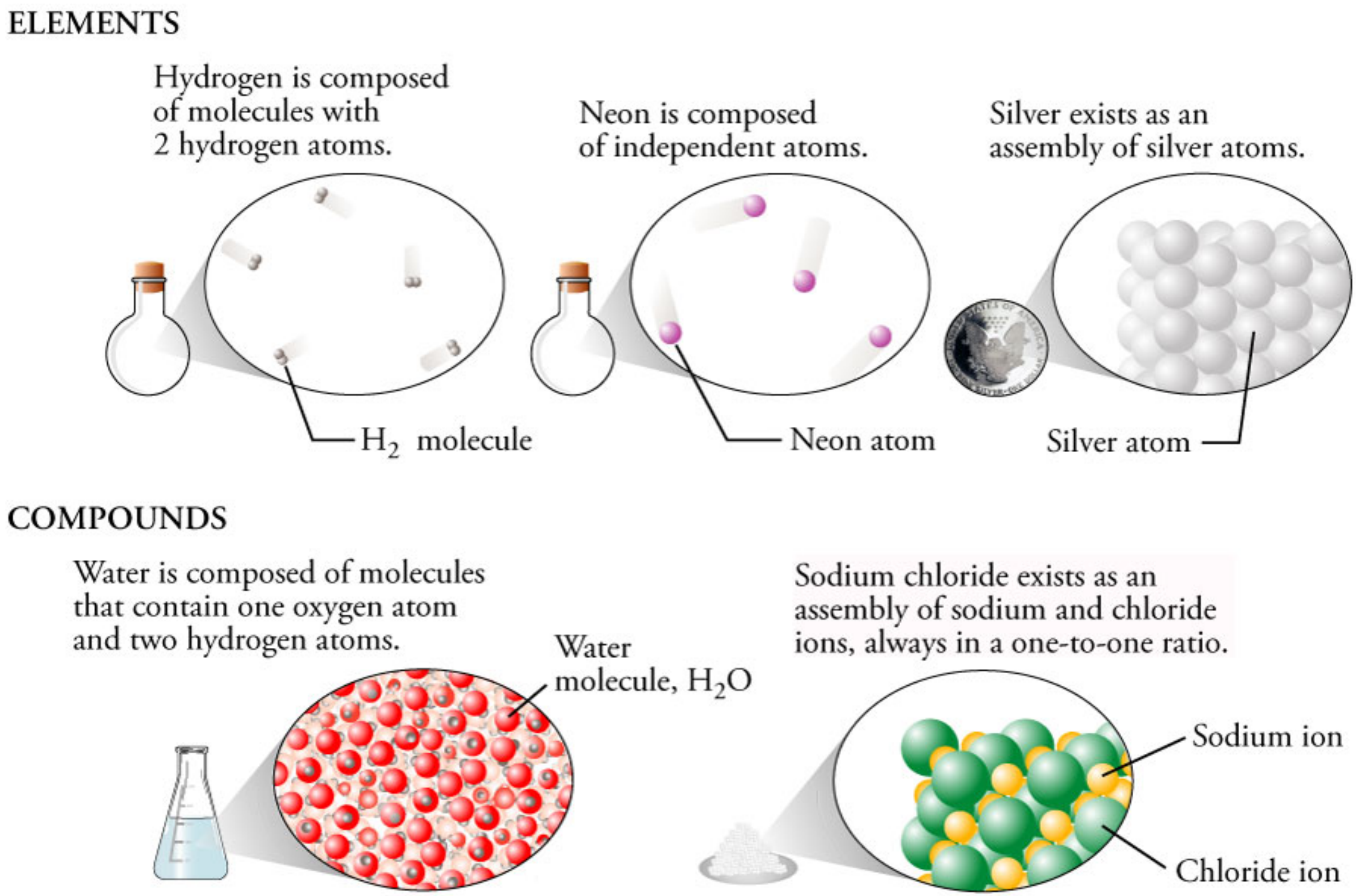
Elements, Compounds, and Mixtures

- **Element:** A substance that cannot be chemically converted into simpler substances; a substance in which all of the atoms have the same number of protons and therefore the same chemical characteristics.
- **Compound:** A substance that contains two or more elements, the atoms of these elements always combining in the same whole-number ratio.
- **Mixture:** A sample of matter that contains two or more pure substances (elements and compounds) and has variable composition.

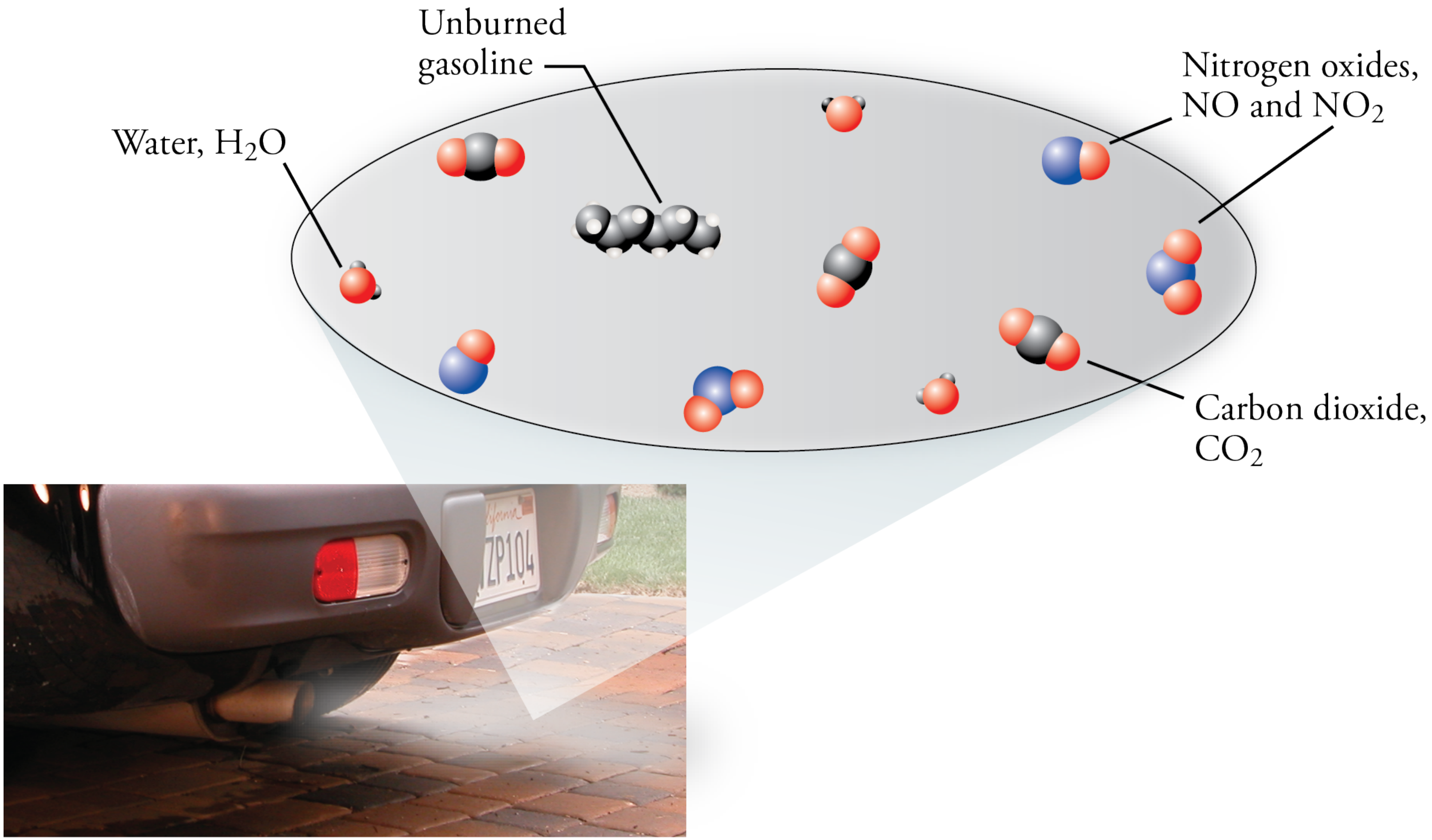
Classification of Matter



Elements and Compounds



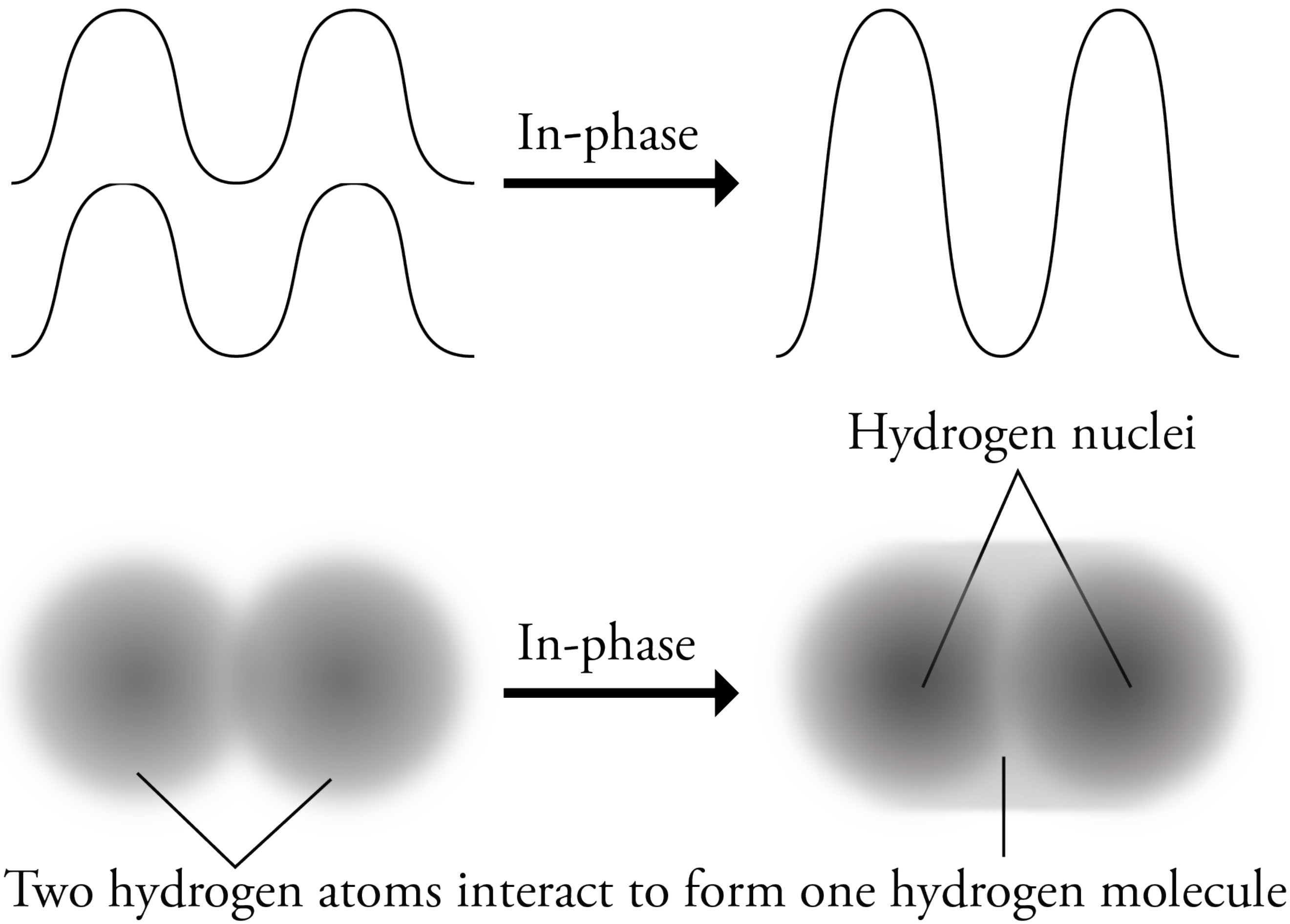
Exhaust – a Mixture



Particle and Wave Nature

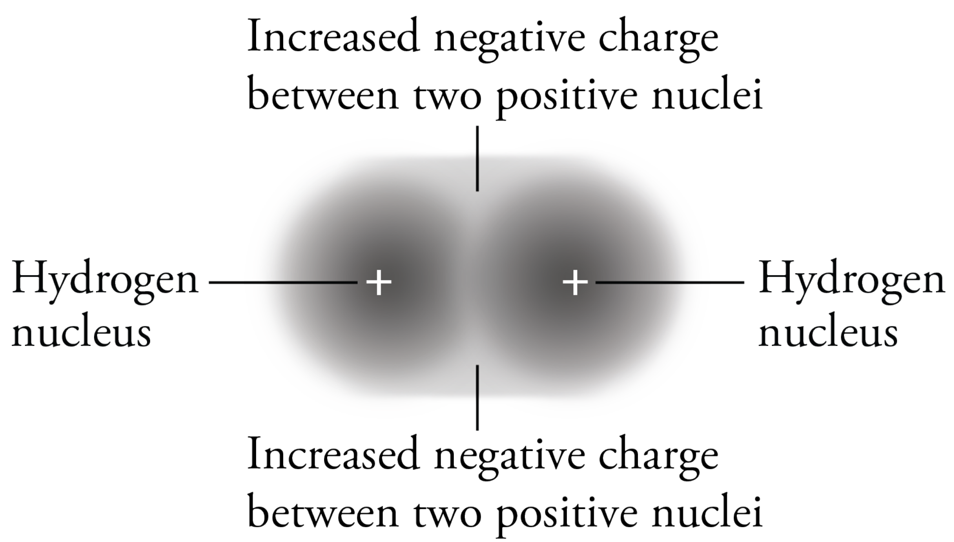
- All matter has both particle and wave character.
- The less massive the particle, the more important its wave character.
- The electron has a very low mass, low enough to have significant wave character.

Covalent Bond Formation



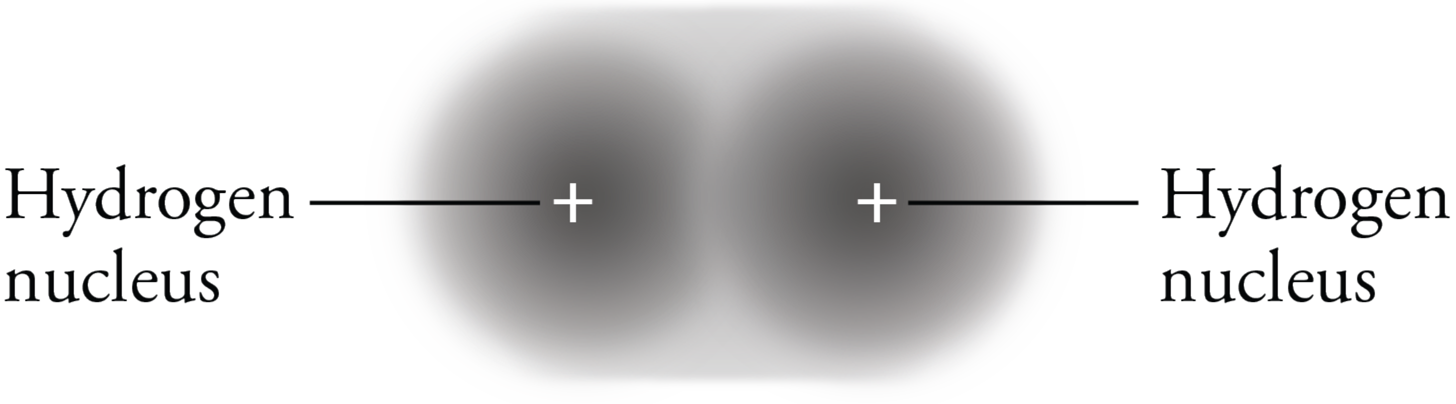
Covalent Bond Formation

- Increased negative charge between the two positive nuclei leads to increased +/- attraction and holds the atoms together.
- **Covalent bond** = a link between atoms due to the sharing of two electrons



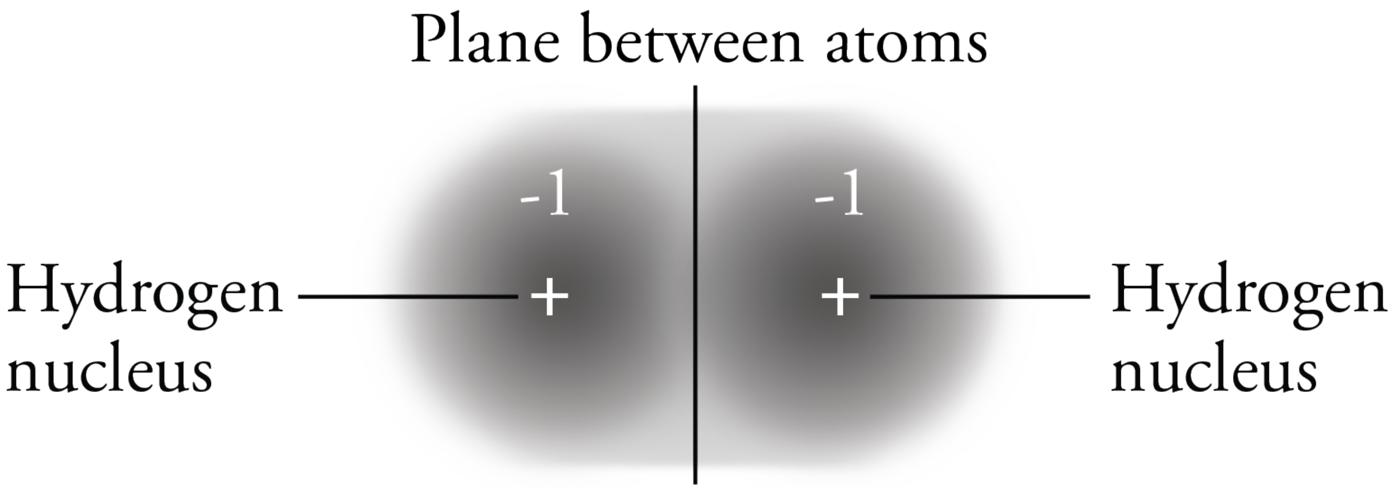
Molecule

- **Molecule** = an uncharged collection of atoms held together by covalent bonds.
- Two hydrogen atoms combine to form a hydrogen molecule, which is described with the formula H_2.



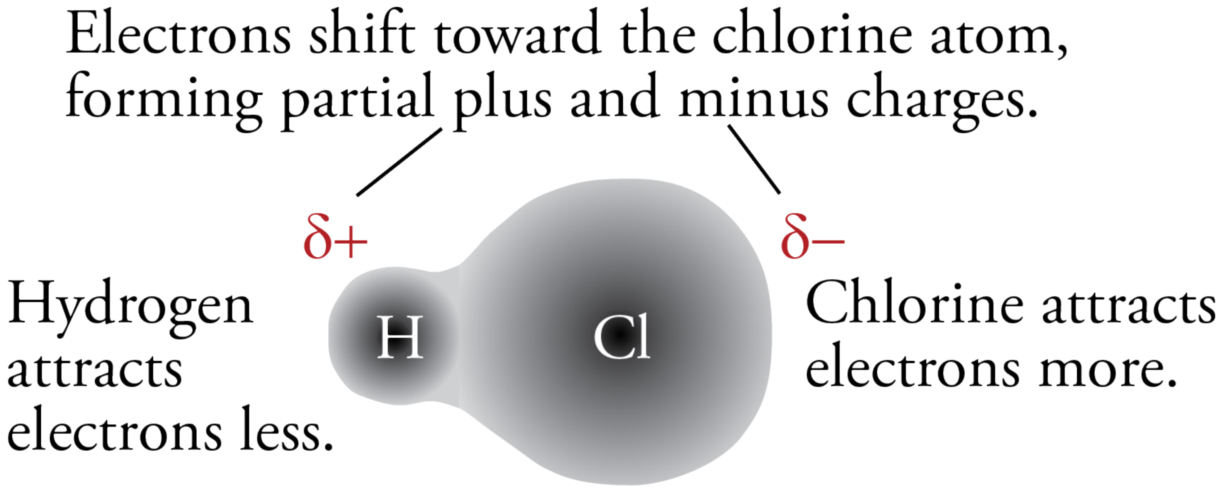
Nonpolar Covalent Bond

- If the electrons are shared equally, there is a even distribution of the negative charge for the electrons in the bond, so there is no partial charges on the atoms. The bond is called a ***nonpolar covalent bond***.



Polar Covalent Bond

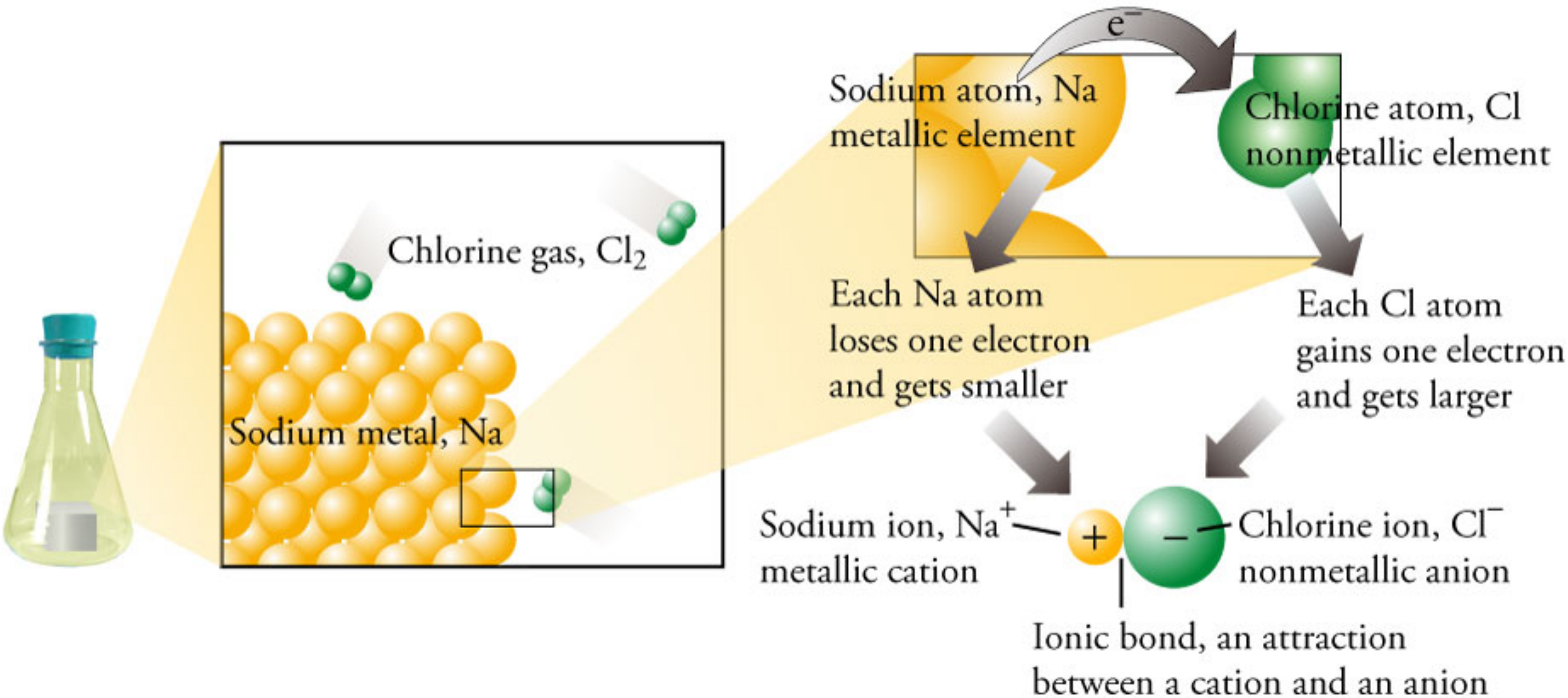
- If one atom in the bond attracts electrons more than the other atom, the electron negative charge shifts to that atom giving it a partial negative charge. The other atom loses negative charge giving it a partial positive charge. The bond is called a ***polar covalent bond***.



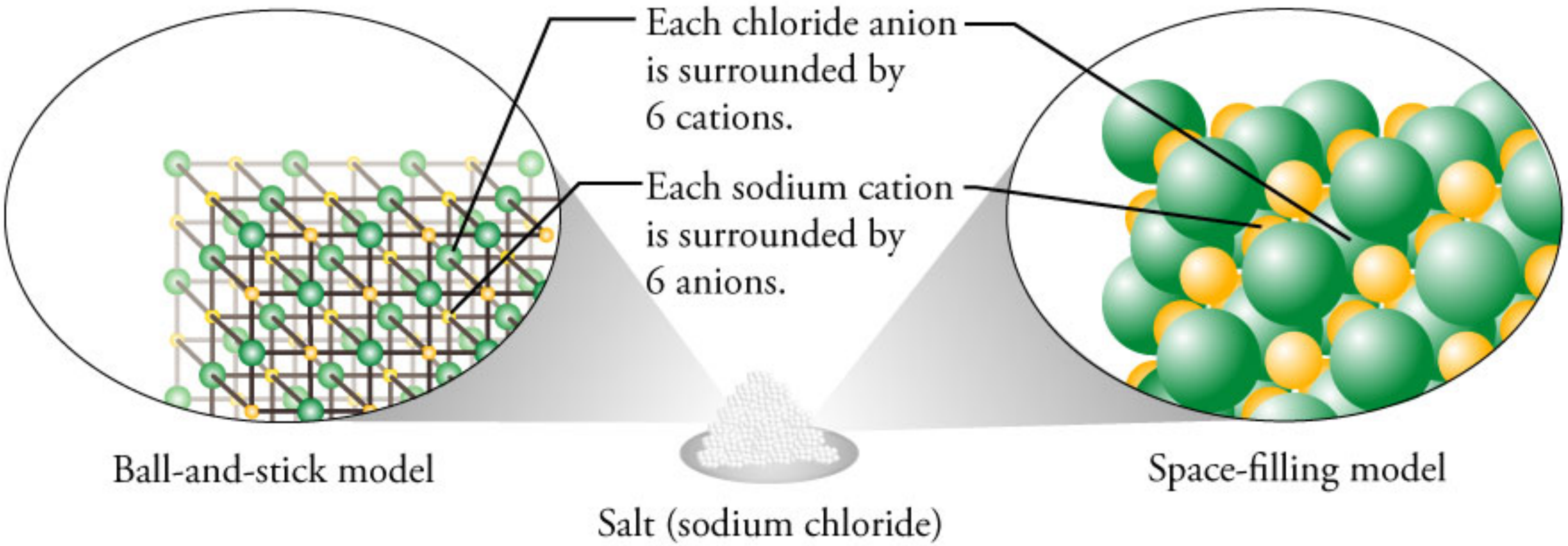
Ionic Bond

- The attraction between cation and anion.
- Atoms of nonmetallic elements often attract electrons so much more strongly than atoms of metallic elements that one or more electrons are transferred from the metallic atom (forming a positively charged particle or ***cation***), to the nonmetallic atom (forming a negatively charged particle or ***anion***).
- For example, an uncharged chlorine atom can pull one electron from an uncharged sodium atom, yielding Cl^- and Na^+.

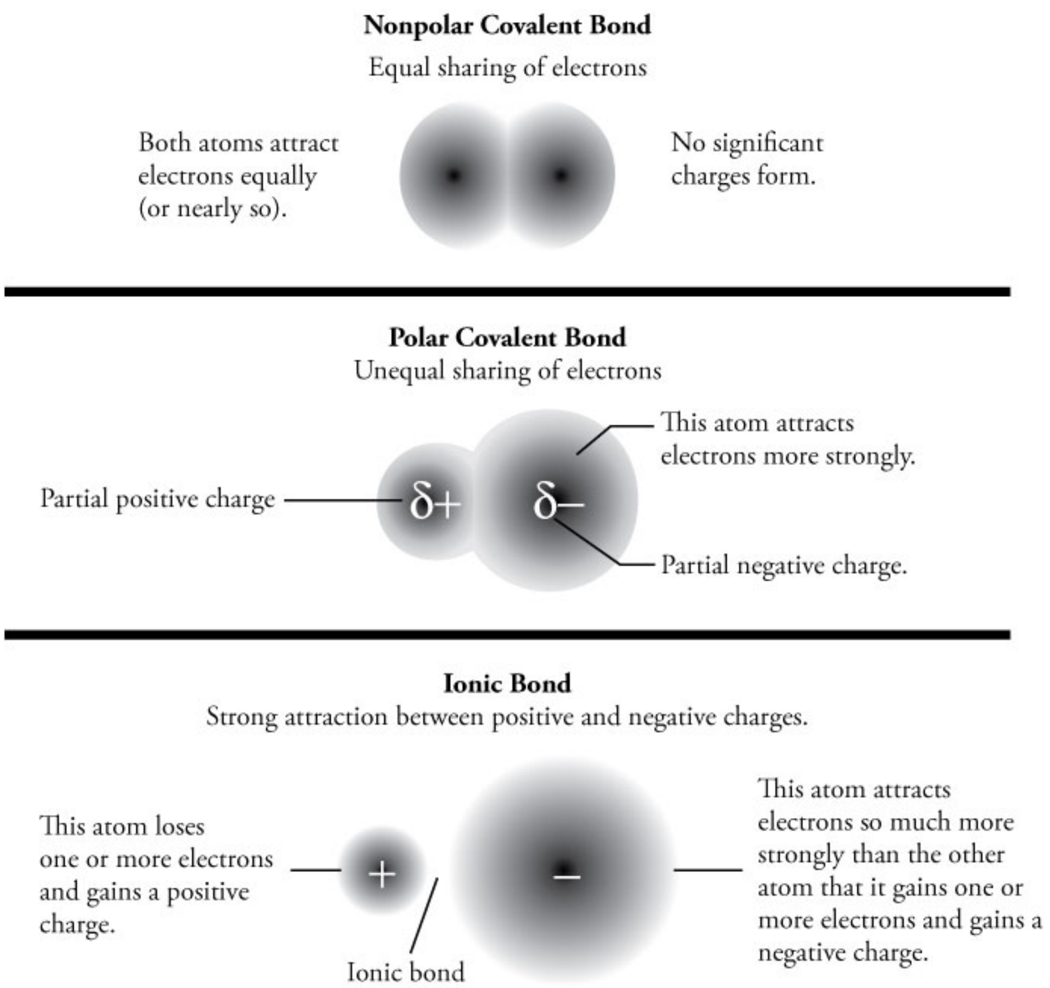
Ionic Bond Formation



Sodium Chloride, NaCl, Structure



Bond Types

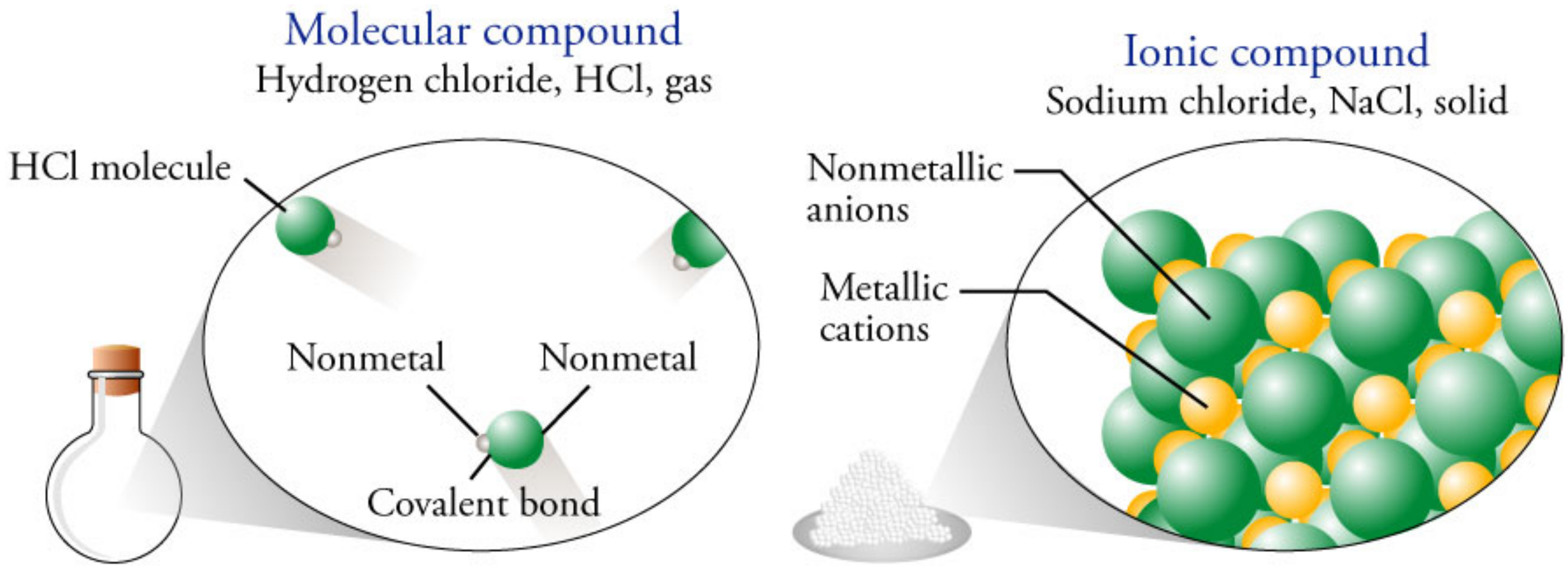


Nonpolar Covalent Bond
Equal sharing of electrons
Both atoms attract electrons equally (or nearly so).
No significant charges form.
Polar Covalent Bond
Unequal sharing of electrons
This atom attracts electrons more strongly.
Partial positive charge
δ+
δ–
Partial negative charge.
Ionic Bond
Strong attraction between positive and negative charges.
This atom loses one or more electrons and gains a positive charge.
+
–
This atom attracts electrons so much more strongly than the other atom that it gains one or more electrons and gains a negative charge.
Ionic bond

Types of Compounds

- All nonmetallic atoms usually leads to all covalent bonds, which from molecules. These compounds are called ***molecular compounds***.
- Metal-nonmetal combinations usually lead to ionic bonds and ***ionic compounds***.

Classification of Compounds



Summary

- **Nonmetal-nonmetal** combinations (e.g. HCl)
  - Covalent bonds
  - Molecules
  - Molecular Compound
- **Metal-nonmetal** combinations (e.g. NaCl)
  - Probably ionic bonds
  - Alternating cations and anions in crystal structure
  - Ionic compound

Cations and Anions

- Atoms of the metallic elements have relatively weak attractions for their electrons, so they tend to lose electrons and form monatomic cations (cations composed of one atom, such as Na^+).
- Atoms of the nonmetallic elements have relatively strong attractions for electrons, so they tend to gain electrons and form monatomic anions (anions composed of one atom, such as Cl^-). .
- Therefore, when metallic and nonmetallic atoms combine, they usually form ions and ionic bonds.

Predicting Ion Charges

- Noble gas atoms are very stable, so when the nonmetallic atoms form anions, they gain enough electrons to get the same number of electrons as the nearest larger noble gas atom.
- When the aluminum and the metallic atoms in Groups 1, 2, and 3 form cations, they lose enough electrons to get the same number of electrons as the nearest smaller noble gas atom.

The Making of an Anion

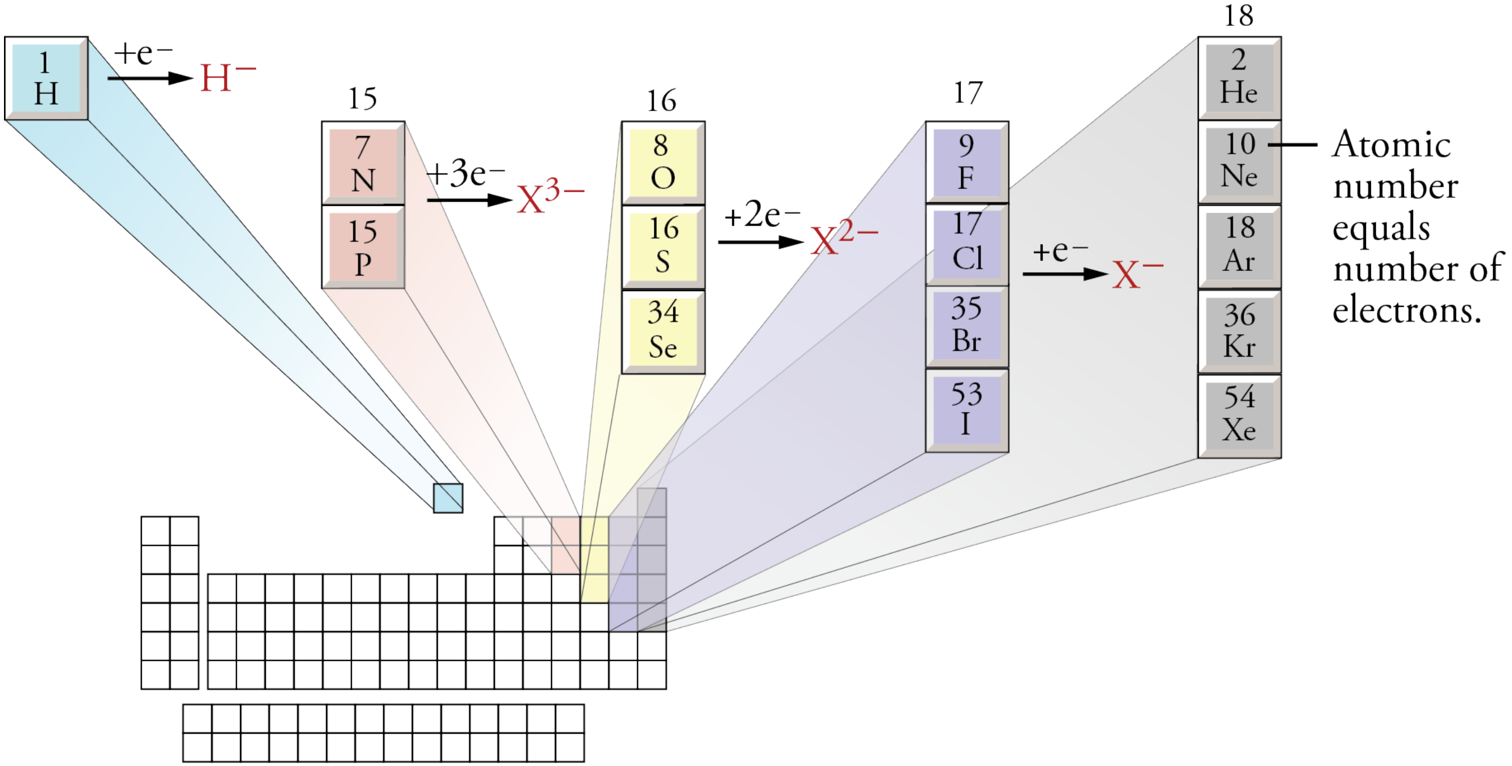
When a hydrogen atom gains one electron,

or when an atom in group 15 gains three electrons,

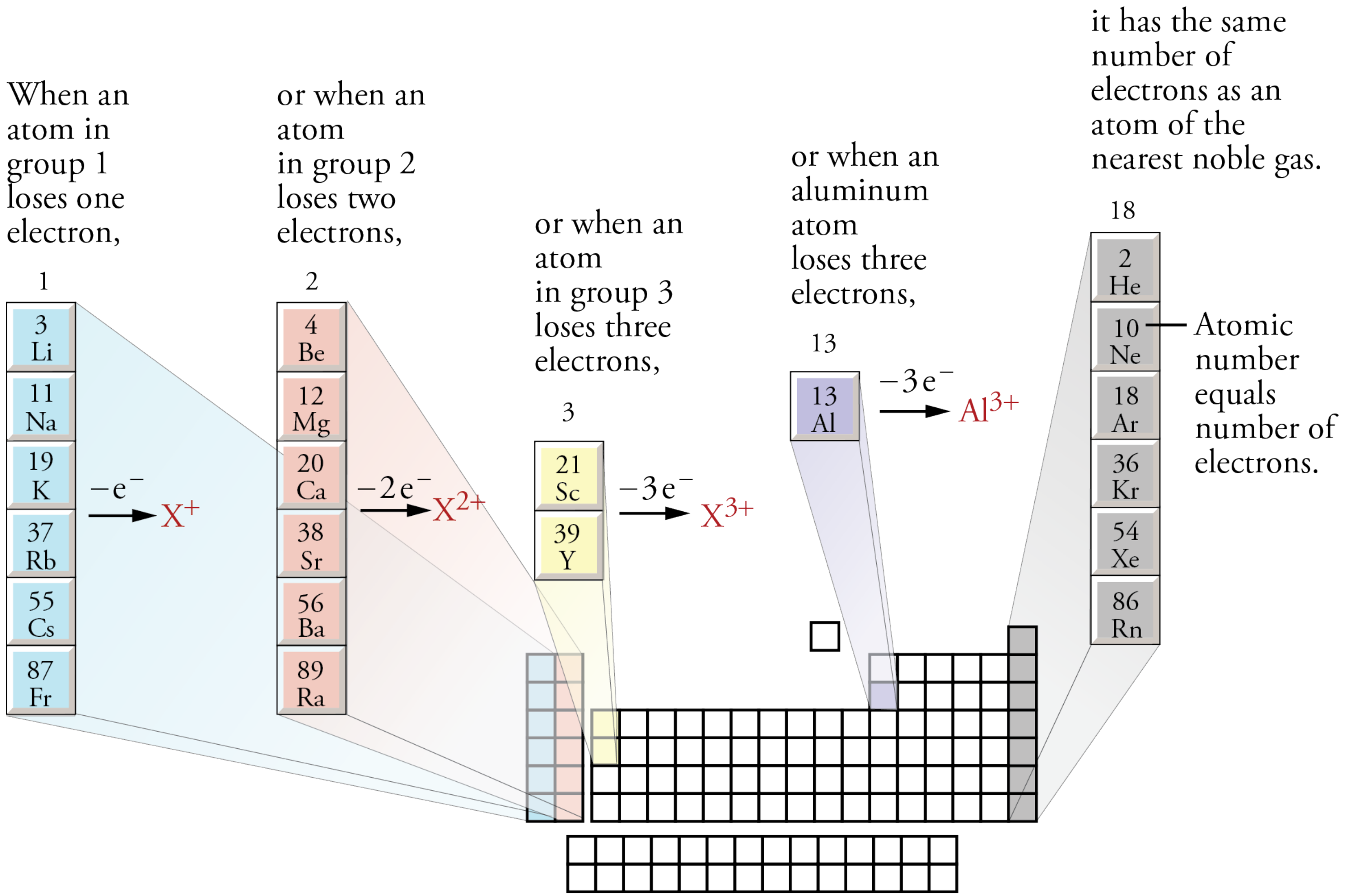
or when an atom in group 16 gains two electrons,

or when an atom in group 17 gains one electron,

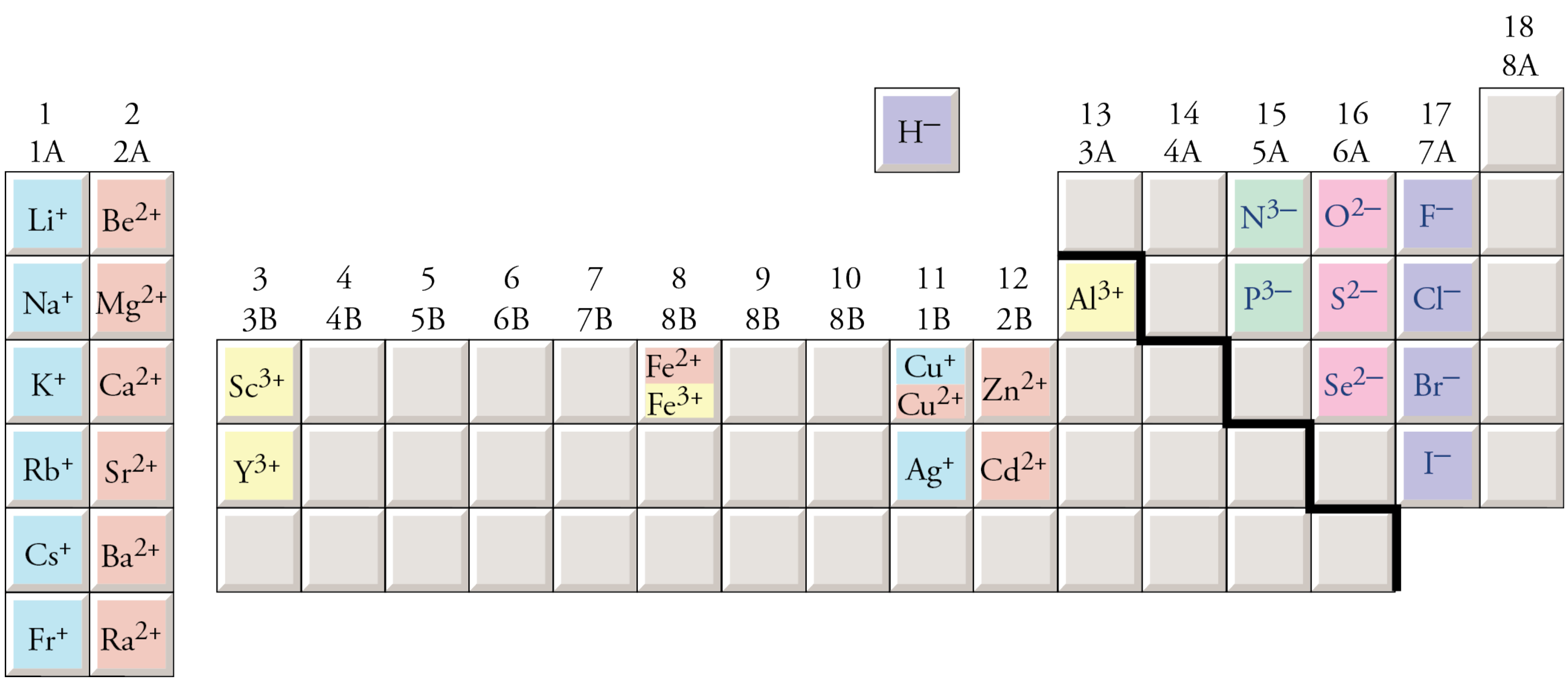
it has the same number of electrons as an atom of the nearest noble gas.



The Making of a Cation



Monatomic Ions



Monatomic Ion Names

- Monatomic Cations
  - (name of metal)
    - Groups 1, 2, and 3 metals
    - Al^{3+}, Zn^{2+}, Cd^{2+}, Ag^{+}
  - (name of metal)(Roman numeral)
    - All metallic cations not mentioned above
- Monatomic Anions
  - (root of nonmetal name)ide

Monatomic Anions

hydride, H^-

nitride, N^{3-}

phosphide, P^{3-}

oxide, O^{2-}

sulfide, S^{2-}

selenide, Se^{2-}

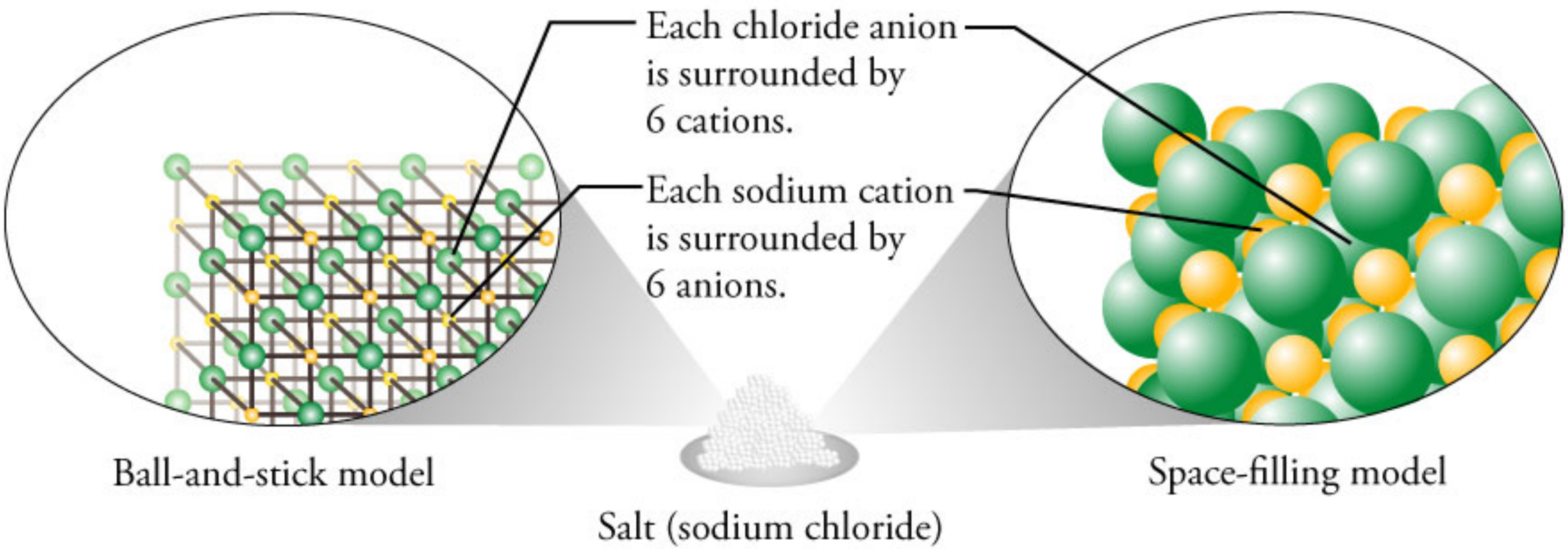
fluoride, F^-

chloride, Cl^-

bromide, Br^-

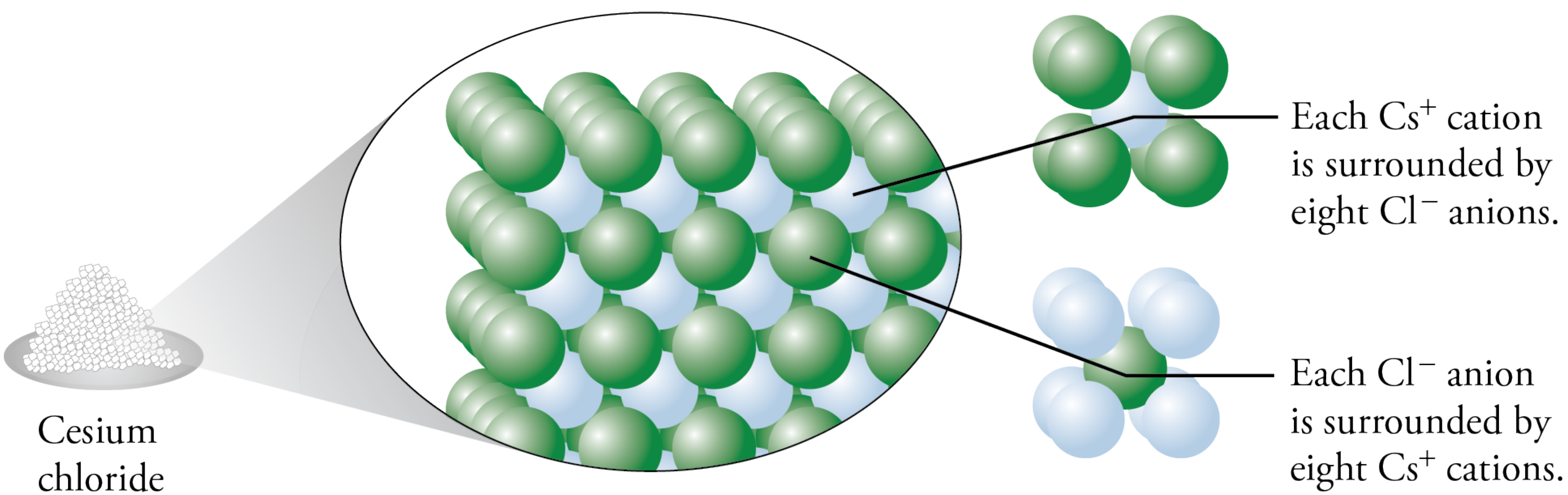
iodide, I^-

Sodium Chloride, NaCl, Structure



Cesium chloride, CsCl, Structure

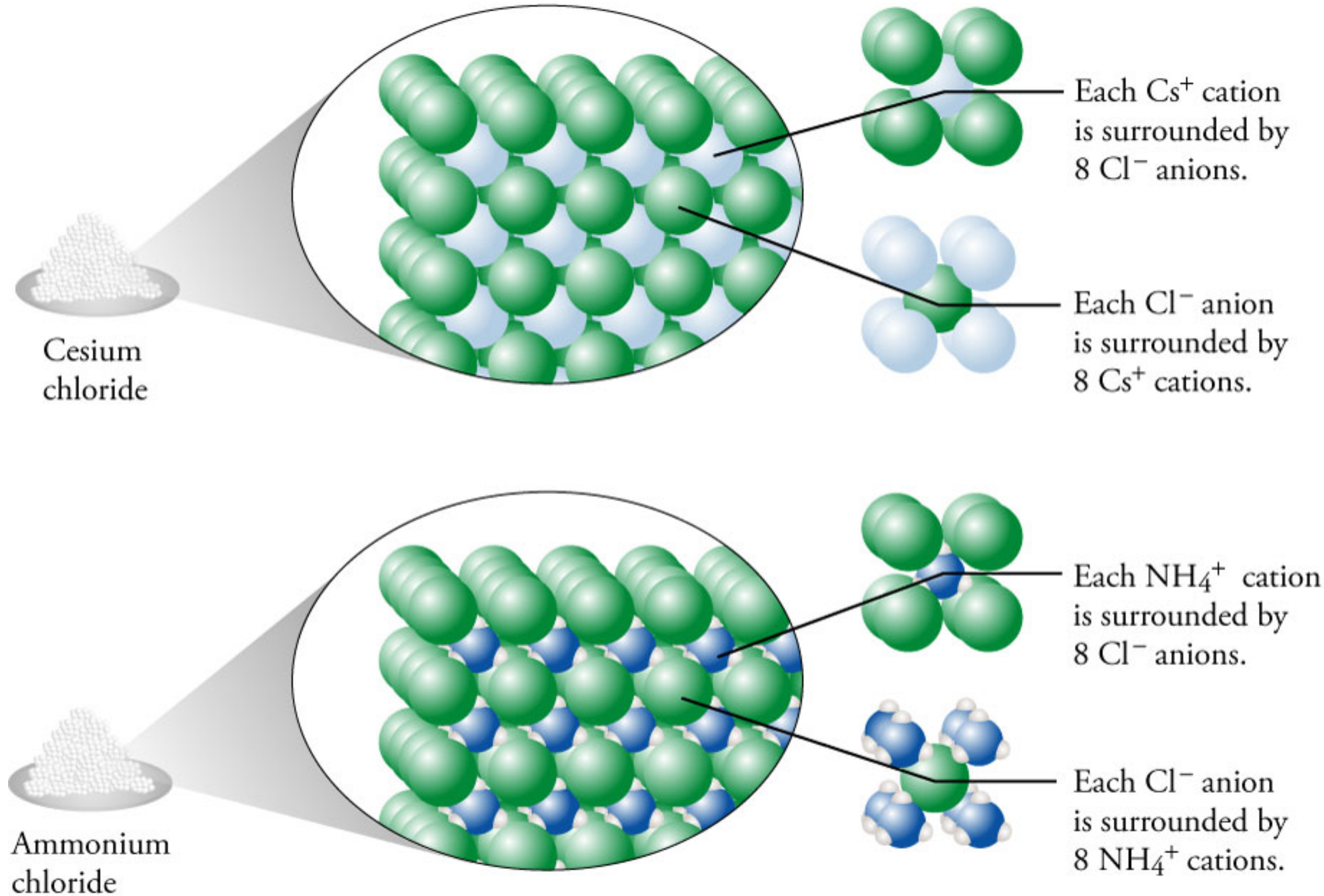
- Because the cesium ions are larger than sodium ions, there is room for eight chloride ions around each cesium ion and eight cesium ions around each chloride ion.



Polyatomic Ions

- Some anions and cations contain more than one atom.
- **Polyatomic ion** = a charge collection of atoms held together by covalent bonds
- For example, it is possible for a nitrogen atom to form covalent bonds to four hydrogen atoms, but to make this possible the nitrogen atom has to lose an electron, giving the collection of atoms a plus one charge. This will be explained in more detail in a later lesson. This collection of atoms with the formula NH_4^+ is called the ammonium ion.

CsCl and NH_4Cl structure



Polyatomic Ions

Ion	Name	Ion	Name
NH_4^+	ammonium	NO_3^-	nitrate
OH^-	hydroxide	SO_4^{2-}	sulfate
CO_3^{2-}	carbonate	$C_2H_3O_2^-$	acetate
PO_4^{3-}	phosphate		

Models – Advantages and Disadvantages (1)

- They help us to *visualize*, *explain*, and *predict* chemical changes.
- Because a model is a *simplified* version of what we think is true, the processes it depicts are sometimes described using the phrase *as if*. When you read, "It is as if an electron were promoted from one orbital to another," the phrase is a reminder that we do not necessarily think this is what really happens. We merely find it useful to talk about the process as if this is the way it happens.

Models – Advantages and Disadvantages (2)

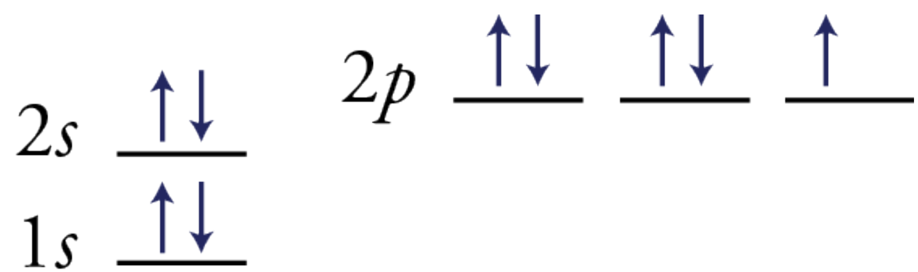
- One characteristic of models is that they ***change with time***. Because our models are simplifications of what we think is real, we are not surprised when they sometimes fail to explain experimental observations. When this happens, the model is altered to fit the new observations.

Assumptions of the Valence-Bond Model

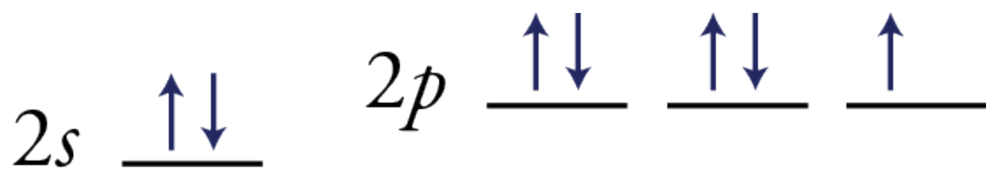
- Only the highest energy electrons participate in bonding.
- Covalent bonds usually form to pair unpaired electrons.

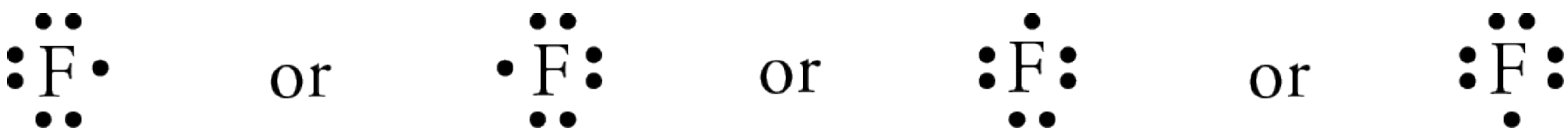
Fluorine

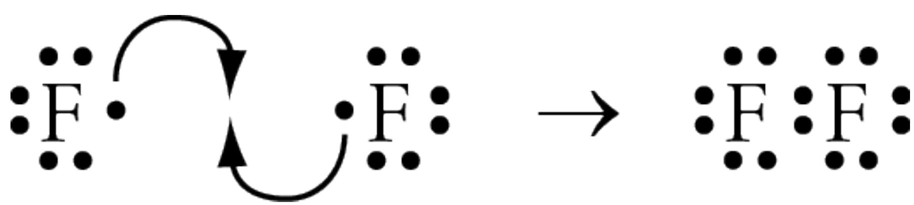
F $1s^2 2s^2 2p^5$



Valence electrons

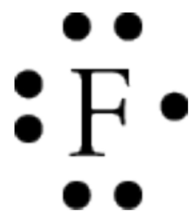






Electron-Dot Symbols and Lewis Structures

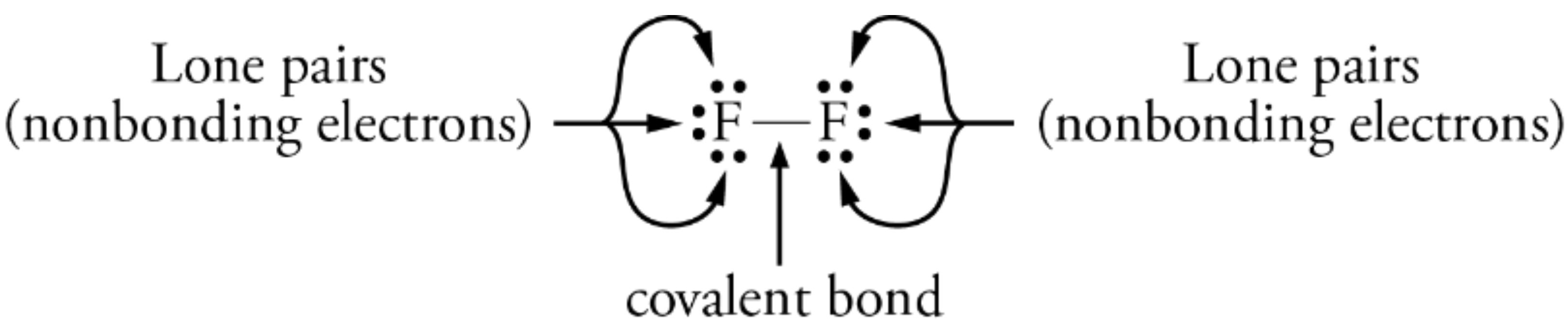
- **Electron-dot symbols** show valence electrons.



- Nonbonding pairs of valence electrons are called ***lone pairs***.

Lewis Structures

- ***Lewis structures*** represent molecules using element symbols, lines for bonds, and dots for lone pairs.



H_2 Formation

- The unpaired electron on a hydrogen atom makes the atom unstable.

$$\mathrm{H}\cdot \qquad\qquad 1s\ \underline{\uparrow}$$

- Two hydrogen atoms combine to form one hydrogen molecule.

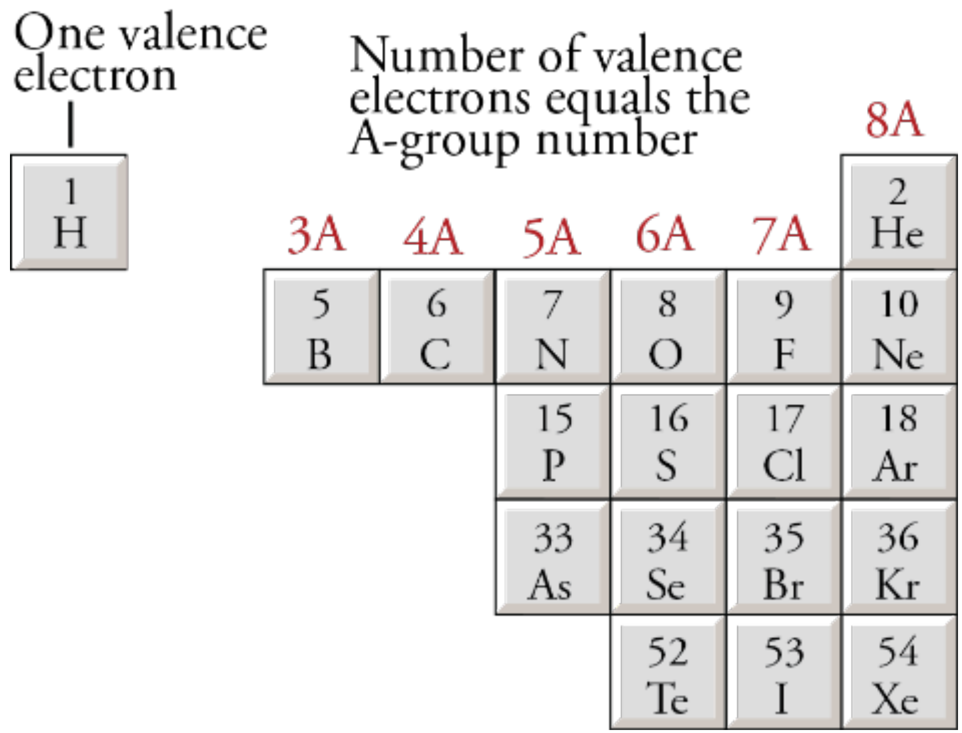
$$\mathrm{H}\cdot \quad \cdot\mathrm{H} \rightarrow \mathrm{H}:\mathrm{H} \quad \text{or} \quad \mathrm{H}—\mathrm{H}$$

Valence Electrons

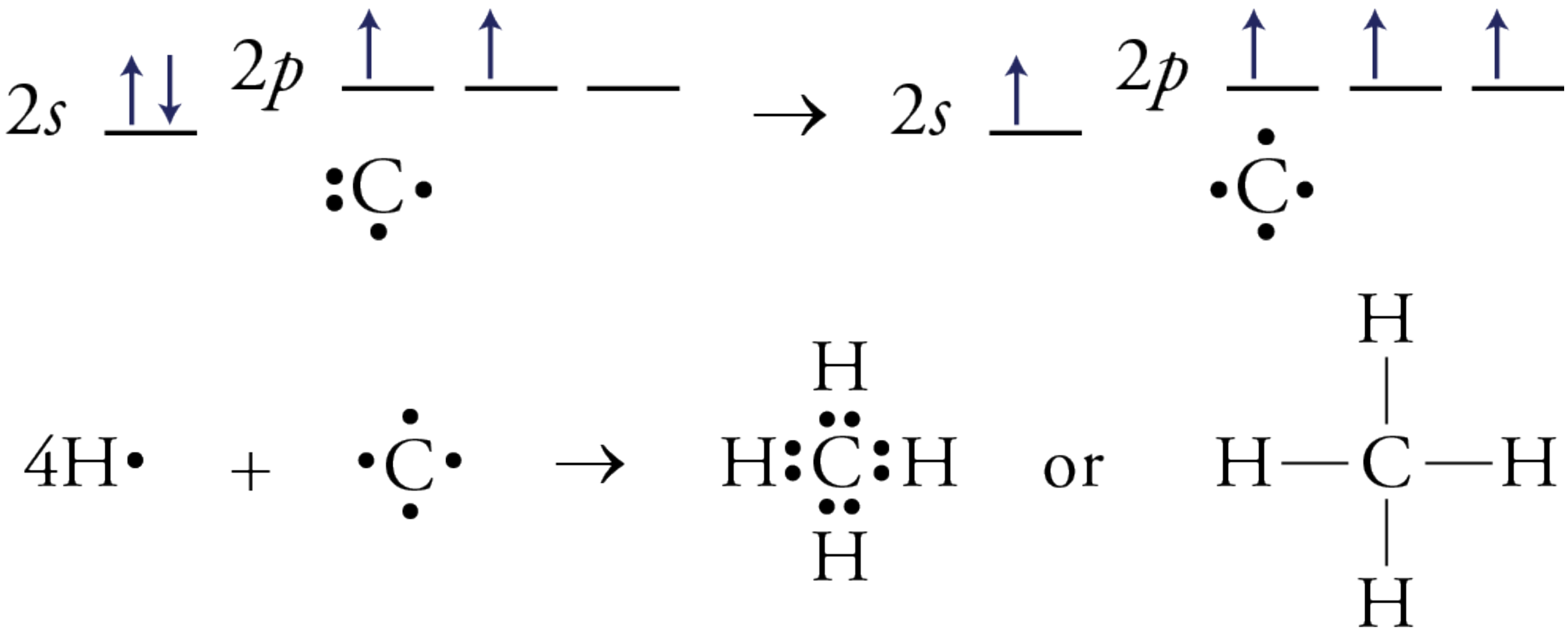
- The valence electrons for each atom are the most important electrons in the formation of chemical bonds.
- The number of valence electrons for the atoms of each element is equal to the element's A-group number on the periodic table.
- Covalent bonds often form to pair unpaired electrons and give the atoms of the elements other than hydrogen and boron eight valence electrons (an octet of valence electrons).

Valence Electrons

- ***Valence electrons*** are the highest-energy *s* and *p* electrons in an atom.



Carbon –
4 bonds



Methane, CH_4

Carbon – Multiple Bonds

$$4\mathrm{H}\cdot \quad + \quad 2\,\cdot\overset{\cdot}{\underset{\cdot}{\mathrm{C}}}\cdot \quad \rightarrow \quad \mathrm{H}\!:\!\underset{\underset{\mathrm{H}}{\cdot\cdot}}{\mathrm{C}}\!::\!\underset{\underset{\mathrm{H}}{\cdot\cdot}}{\mathrm{C}}\!:\!\mathrm{H} \quad \text{or} \quad \mathrm{H}-\underset{\underset{\mathrm{H}}{|}}{\mathrm{C}}=\underset{\underset{\mathrm{H}}{|}}{\mathrm{C}}-\mathrm{H}$$

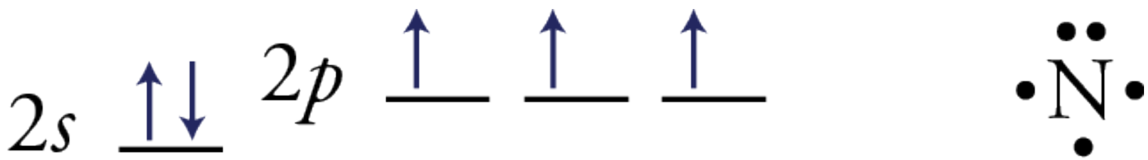
Ethene (ethylene)

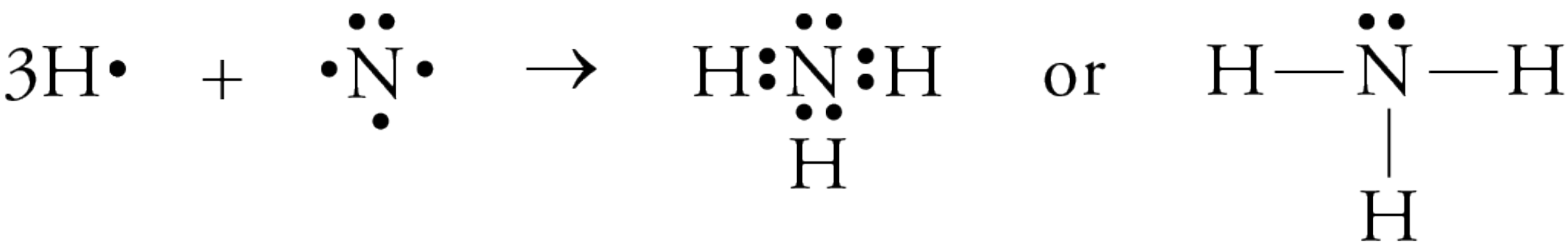
$$2\mathrm{H}\cdot \quad + \quad 2\,\cdot\overset{\cdot}{\underset{\cdot}{\mathrm{C}}}\cdot \quad \rightarrow \quad \mathrm{H}\!:\!\mathrm{C}\!\vdots\vdots\!\mathrm{C}\!:\!\mathrm{H} \quad \text{or} \quad \mathrm{H}-\mathrm{C}\equiv\mathrm{C}-\mathrm{H}$$

Ethyne (acetylene)

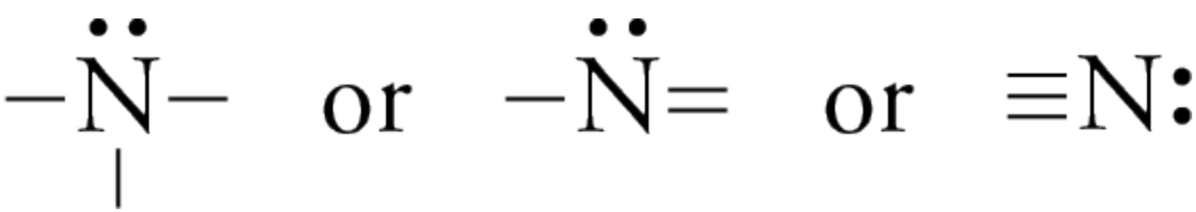
$$-\overset{|}{\underset{|}{\mathrm{C}}}- \quad \text{or} \quad -\underset{|}{\mathrm{C}}= \quad \text{or} \quad =\mathrm{C}= \quad \text{or} \quad -\mathrm{C}\equiv$$

Nitrogen – 3 bonds & 1 lone pair

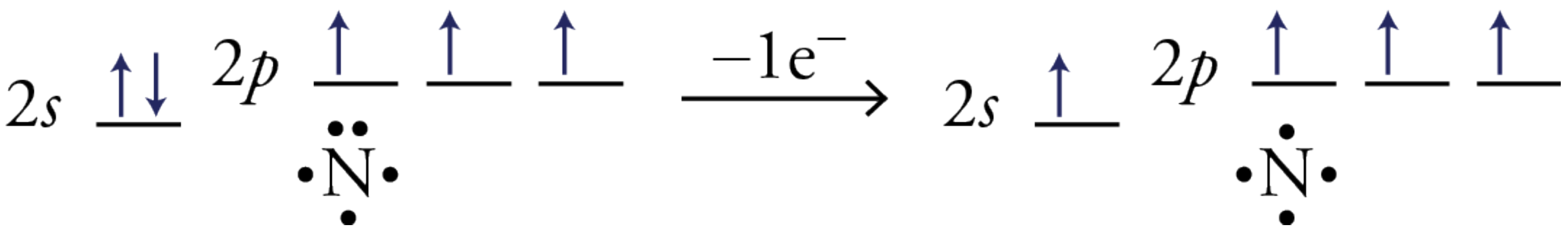


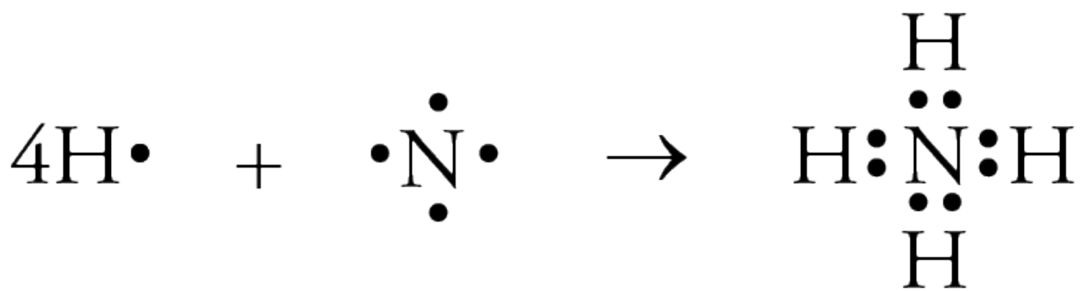


Ammonia, NH_3



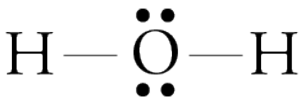
Nitrogen – 4 bonds



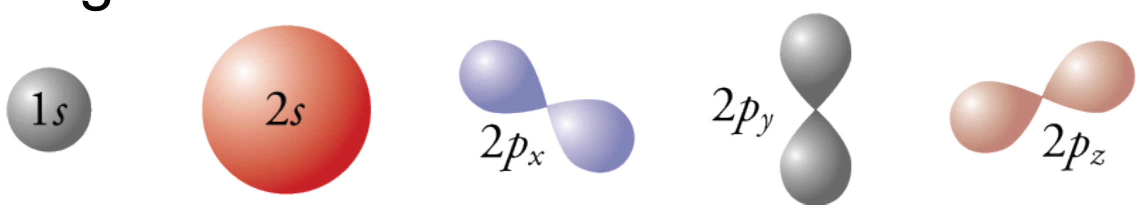


Ammonium, NH_4^+

Where does the following Lewis structure come from?

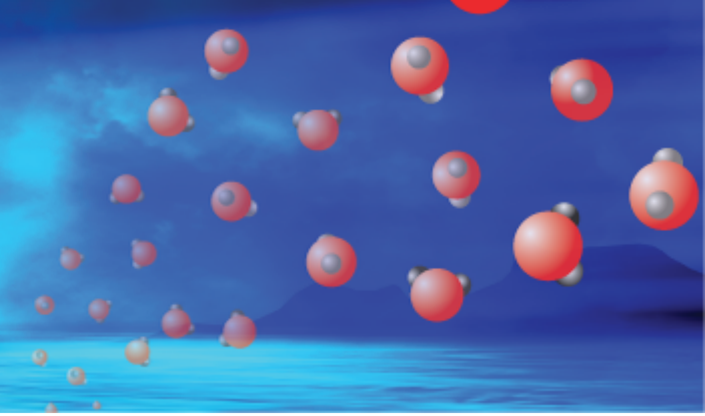


- 3-D wave mathematics for the one electron of hydrogen leads to



- Assume that all other elements have hydrogen-like orbitals.
- Recognize that electrons can have one of only two possible spins.
- Assume that no two electrons in the same atom can be exactly the same.

How Did We Get here?



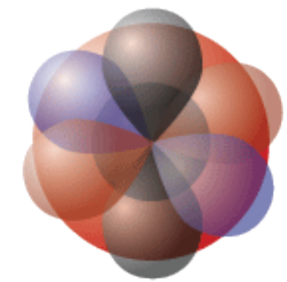
- Leads to orbital diagram of oxygen.

2s ↑↓ 2p ↑↓ ↑ ↑

1s ↑↓

O

$1s^2\ 2s^2\ 2p^4$

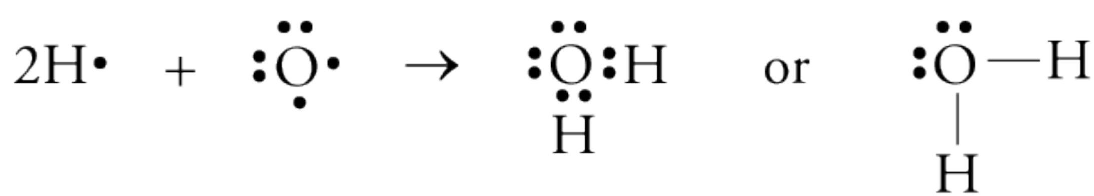


- Assume that only the highest energy electrons participate in bonding.

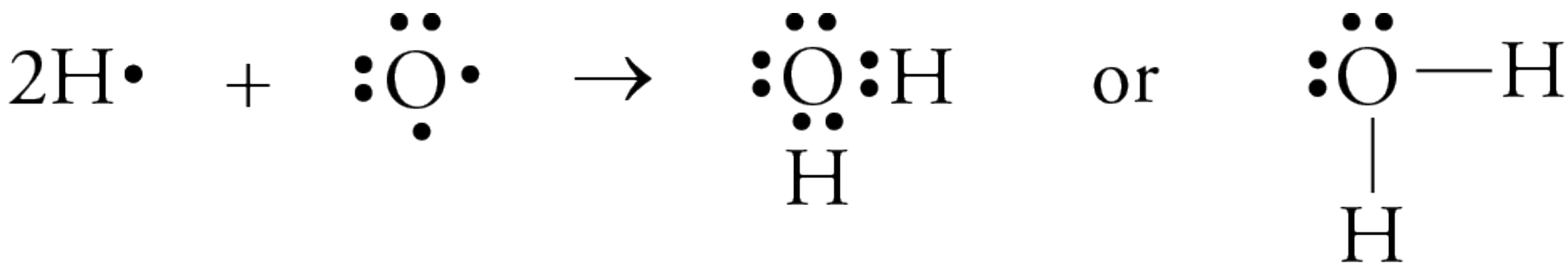
2s ↑↓ 2p ↑↓ ↑ ↑

$2s^2\ 2p^4$

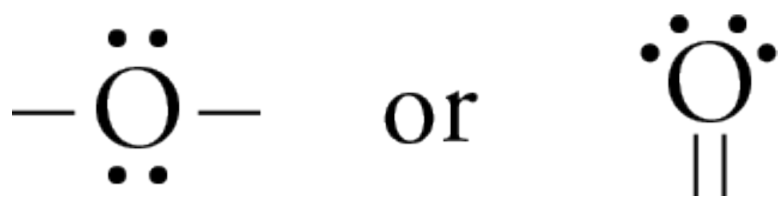
- Assume that covalent bonds form to pair, unpaired electrons.



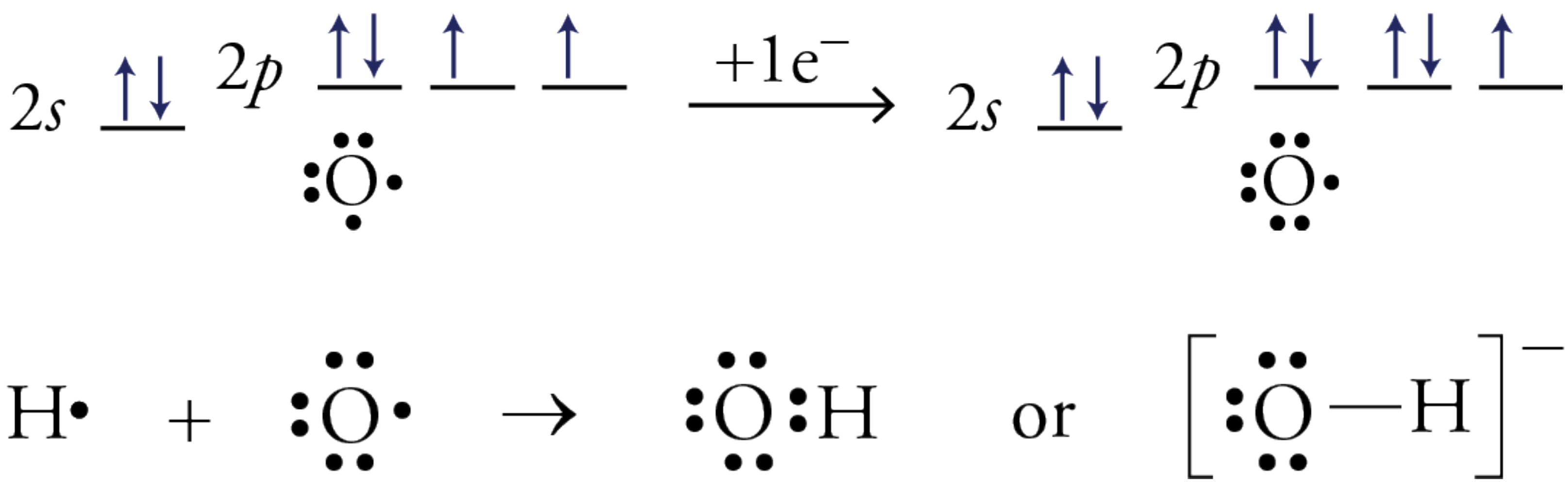
Oxygen – 2 bonds & 2 lone pairs



Water, H_2O



Oxygen – 1 bond & 3 lone pairs

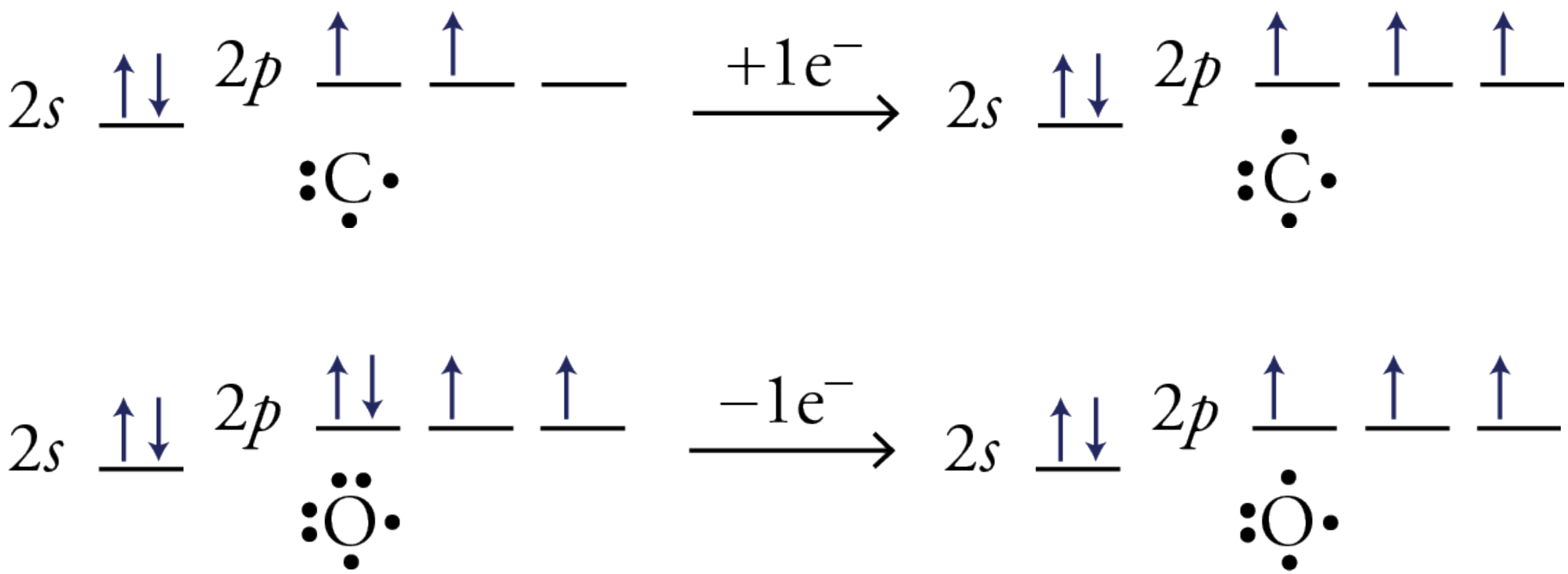


Hydroxide, OH^-

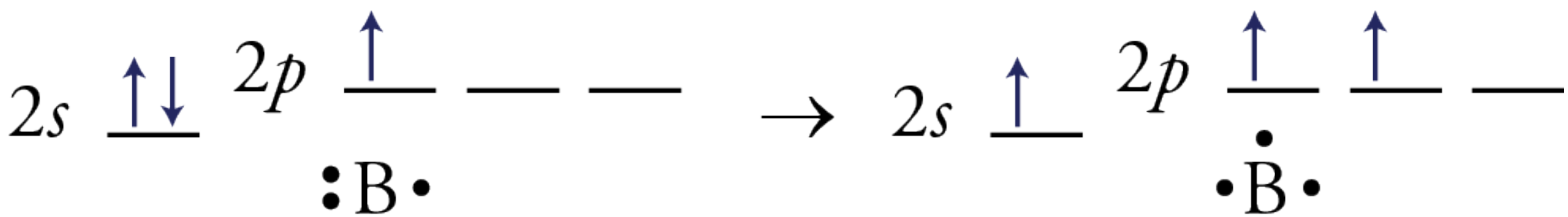
Carbon – 3 bonds & 1 lone pair
Oxygen – 3 bonds & 1 lone pair

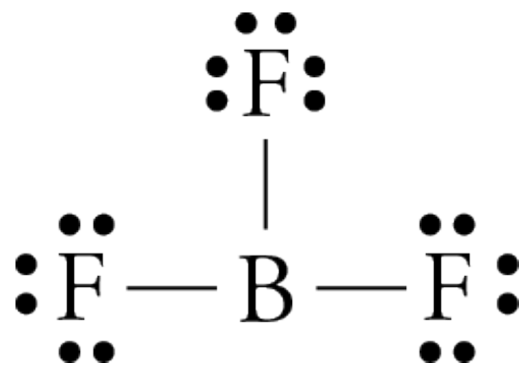
$$:\mathrm{C}\equiv\mathrm{O}:$$

Carbon monoxide, CO



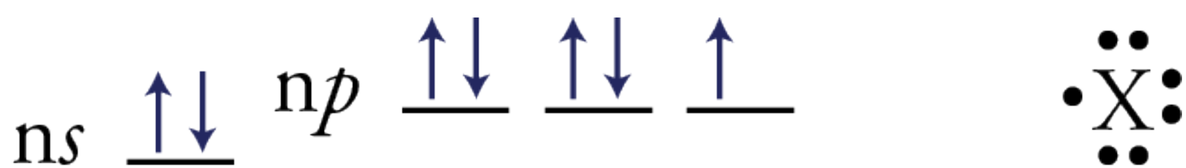
Boron –
3 bonds

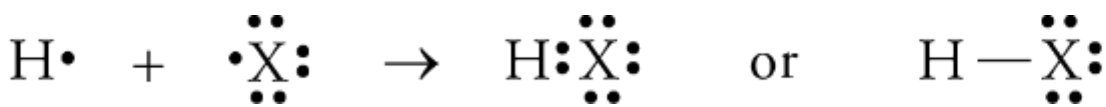




Boron trifluoride, BF_3

Halogens – 1 bond & 3 lone pairs





X = F, Cl, Br, or I

Most Common Bonding Patterns for Nonmetals

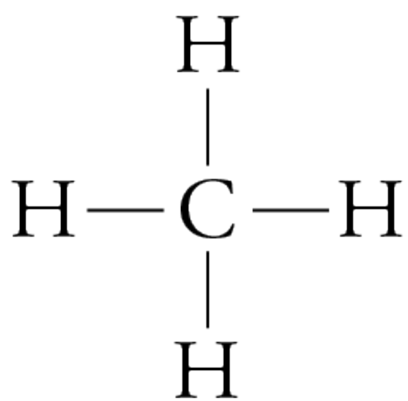
Element	# Bonds	# lone pairs
H	1	0
C	4	0
N, P	3	1
O, S, Se	2	2
F, Cl, Br, I	1	3

Drawing Lewis Structures

- This chapter describes a procedure that allows you to draw Lewis structures for many different molecules.
- Many Lewis structures can be drawn by attempting to give each atom in a molecule its most common bonding pattern.

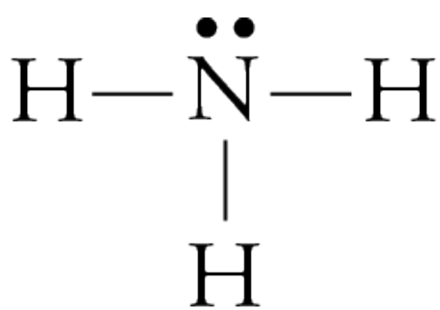
Lewis Structure for Methane, CH_4

- Carbon atoms usually have 4 bonds and no lone pairs.
- Hydrogen atoms have 1 bond and no lone pairs.



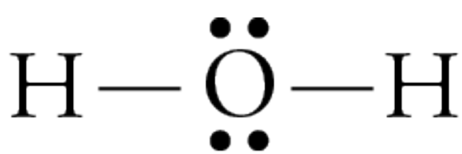
Lewis Structure for Ammonia, NH_3

- Nitrogen atoms usually have 3 bonds and 1 lone pair.
- Hydrogen atoms have 1 bond and no lone pairs.



Lewis Structure for Water, H_2O

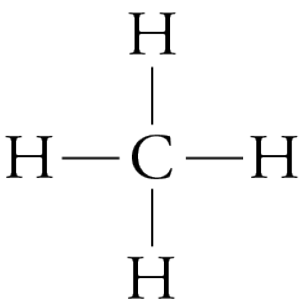
- Oxygen atoms usually have 2 bonds and 2 lone pairs.
- Hydrogen atoms have 1 bond and no lone pairs.

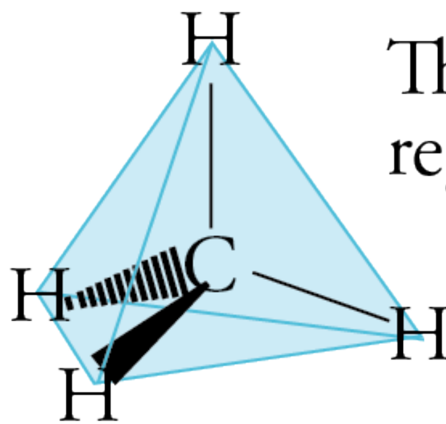


Drawing Reasonable Lewis Structures Skeletons

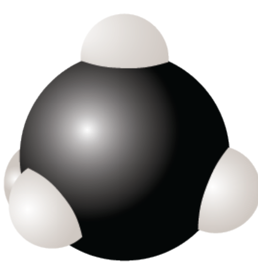
- Try to arrange the atoms to yield the most typical number of bonds for each atom.
- Apply the following guidelines in deciding what element belongs in the center of your structure.
  - Hydrogen and fluorine atoms are never in the center.
  - Oxygen atoms are rarely in the center.
  - The element with the fewest atoms in the formula is often in the center.
  - The atom that is capable of making the most bonds is often in the center.
- Oxygen atoms rarely bond to other oxygen atoms.
- The molecular formula often reflects the molecular structure.
- Carbon atoms commonly bond to other carbon atoms.

Methane, CH_4

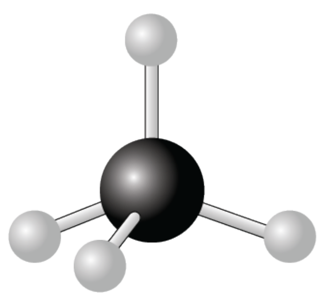




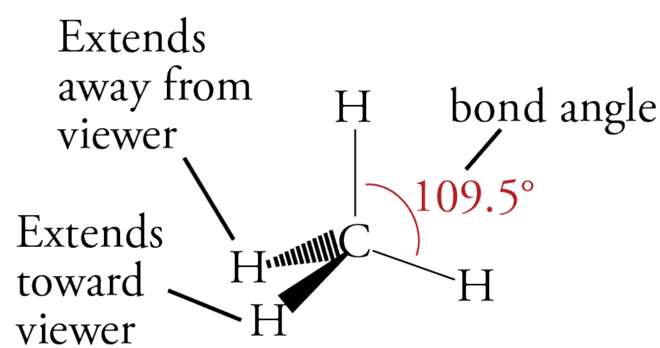
The shaded shape is a regular tetrahedron.



Space-filling model

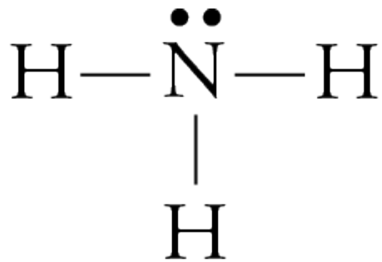


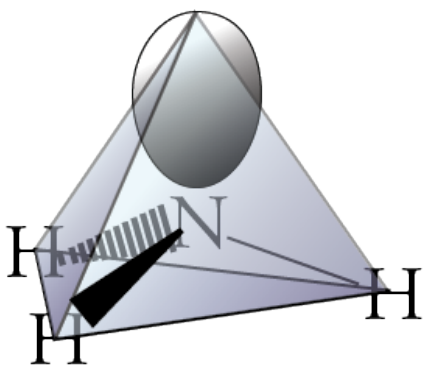
Ball-and-stick model



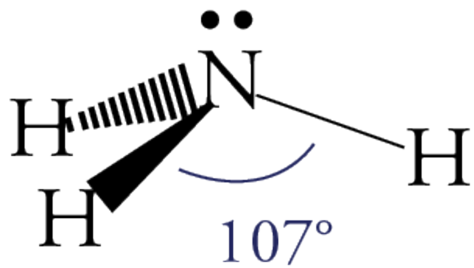
Geometric Sketch

Ammonia, NH_3

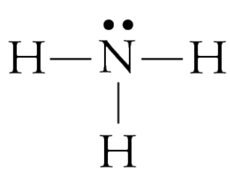




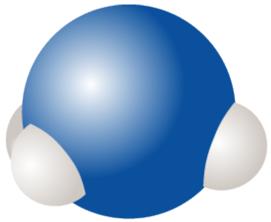
Electron group geometry (tetrahedral)

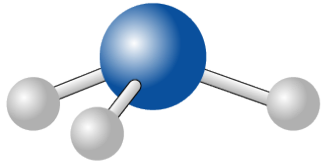


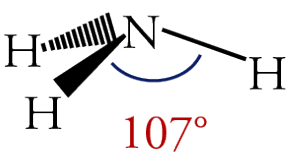
Molecular geometry (trigonal pyramid)



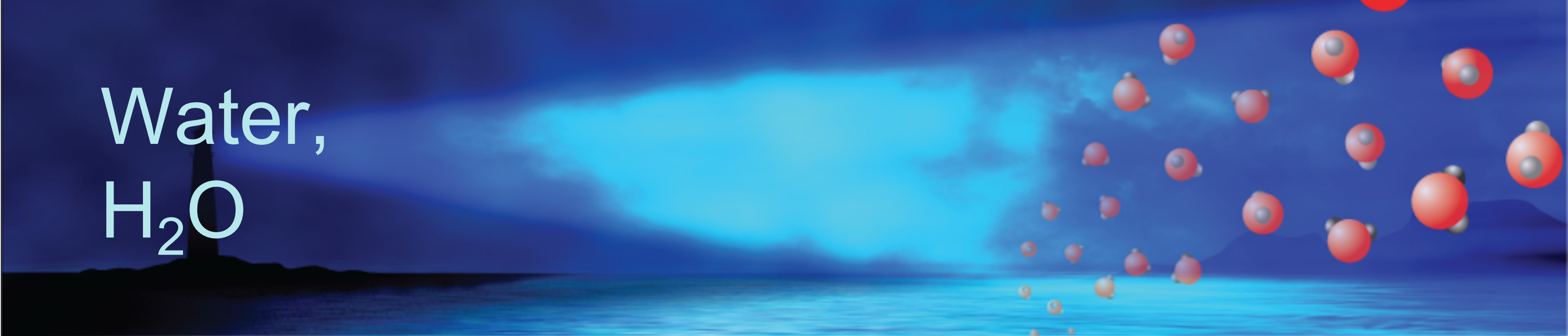
Lewis structure


Space-filling model

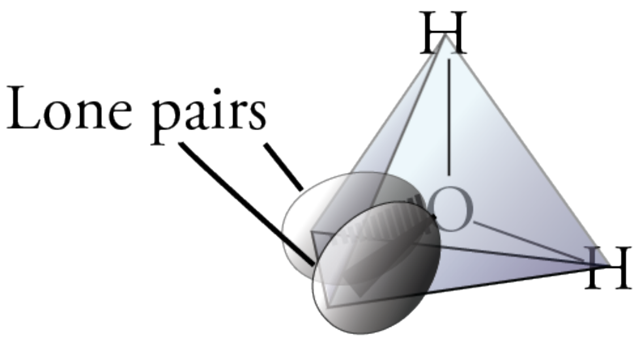

Ball-and-stick model

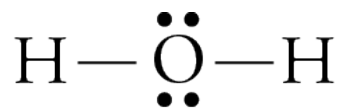


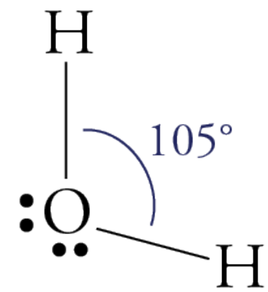
Geometric sketch



Water, H_2O

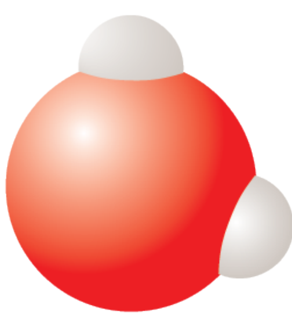


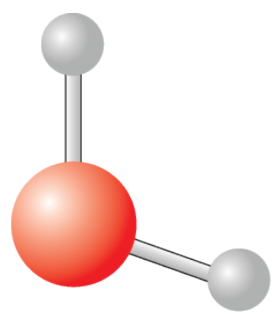


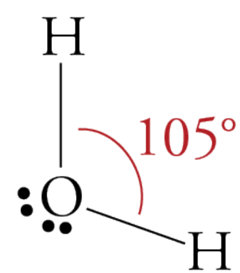


Electron group geometry (tetrahedral)

Molecular geometry (bent)





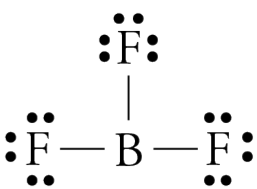


Space-filling model

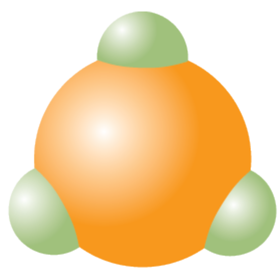
Ball-and-stick model

Geometric Sketch

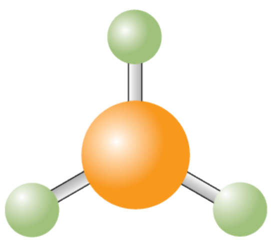
Trigonal Planar Geometry – BF_3



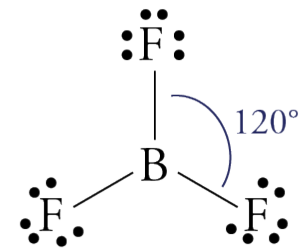
Lewis structure



Space-filling model

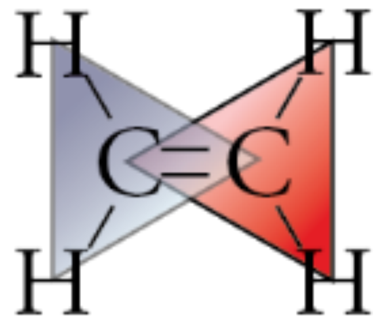


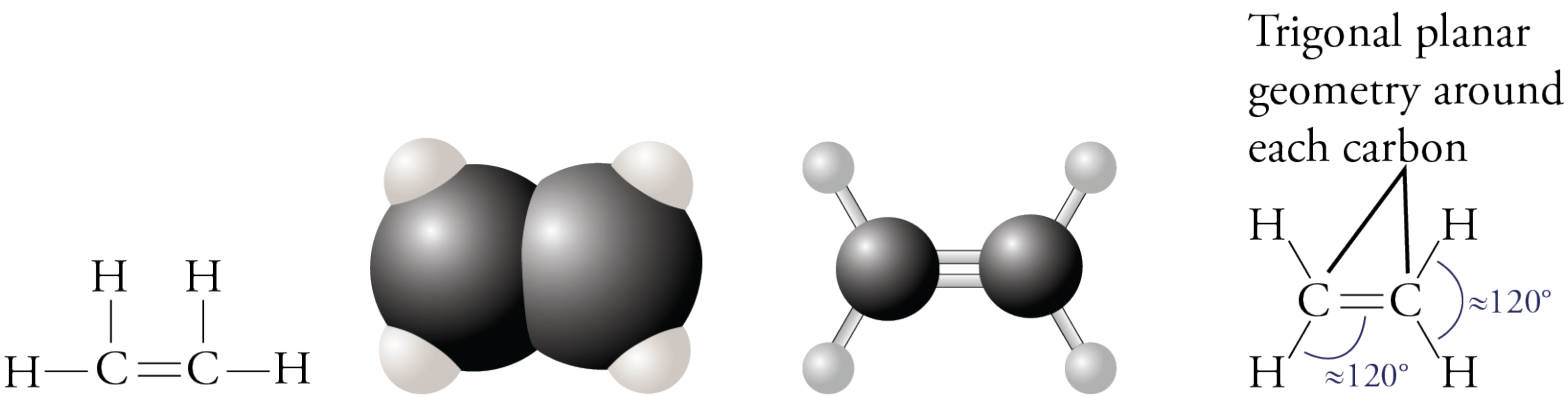
Ball-and-stick model



Geometric Sketch

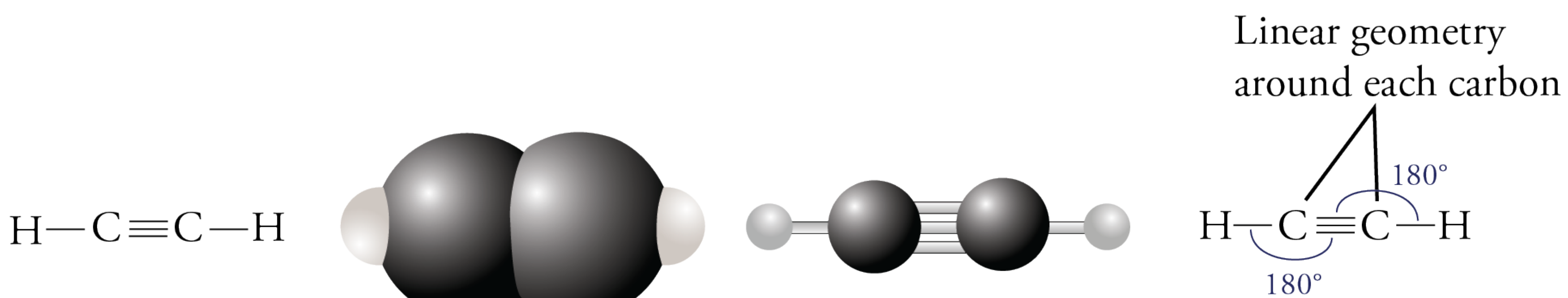
Ethene (ethylene)





Lewis structure Space-filling model Ball-and-stick model Geometric Sketch

Ethyne (acetylene) , C_2H_2



Lewis structure

Space-filling model

Ball-and-stick model

Geometric Sketch

Steps for Molecular Geometry

- **Step 1:** To determine the name of the electron group geometry around each atom that is attached to two or more atoms, count the number of electron groups around each atom and apply the guidelines found on Table 5.2.
- **Step 2:** Use one or more of the geometric sketches shown on Table 5.2 for the geometric sketch of your molecule.

Steps for Molecular Geometry (cont.)

- **Step 3:** To determine the name of the molecular geometry around each atom that has two or more atoms attached to it, count the number of bond groups and lone pairs, and then apply the guidelines found on Table 5.2.