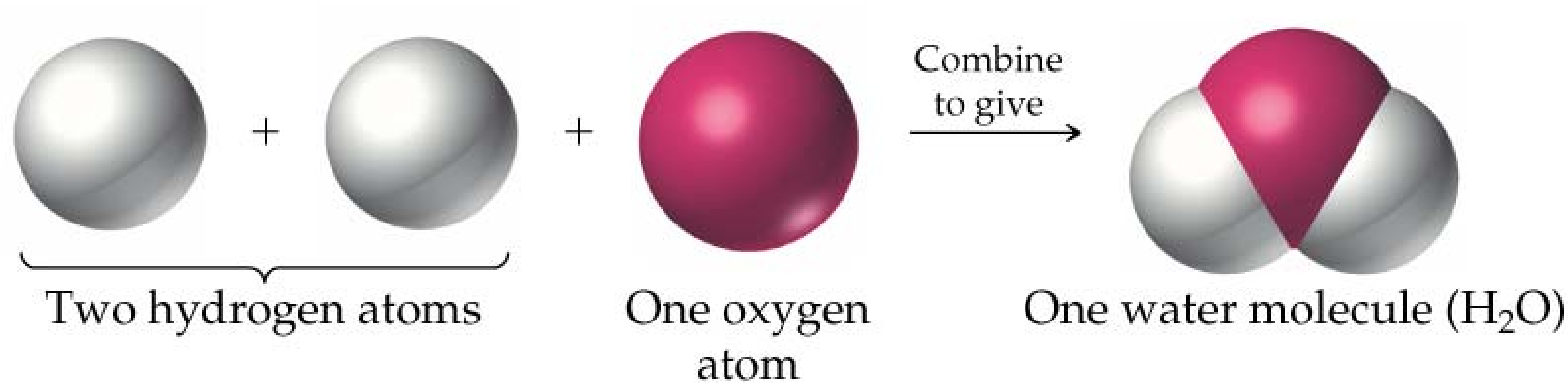
5. 1 Covalent Bonds

- ***Covalent bond***: A bonds formed by sharing electrons between atoms.
- ***Molecule***: A group of atoms held together by covalent bonds.
- The nonmetals near the middle of the periodic table reach an electron octet by sharing an appropriate number of electrons.

- Water molecule results when two hydrogen atoms and one oxygen atom are covalently bonded in a way shown in the following picture:



- When two atoms come together, electrical interactions occur.
- Some of these interactions are repulsive – positively charged nuclei repel each other and the negatively charged electrons repel each other.
- Other interactions are attractive – each nucleus attract electrons of both atoms and electrons in both atoms attract each nucleus.
- Because attractive forces are stronger than the repulsive forces, a covalent bond is formed between the atoms.

A covalent bond between two hydrogen atoms is shown in this picture.

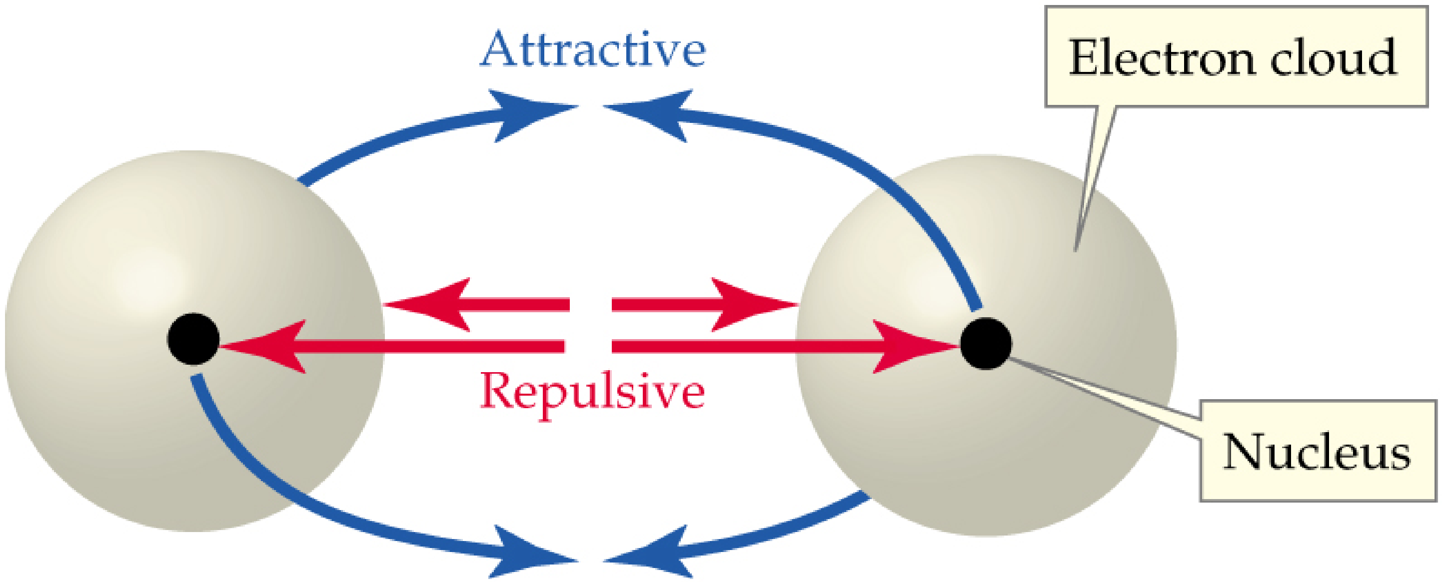
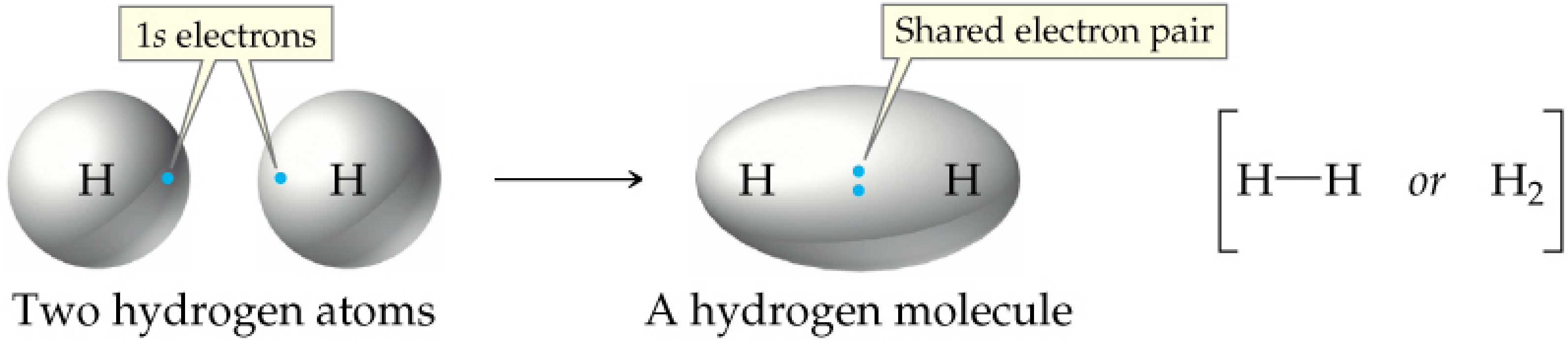
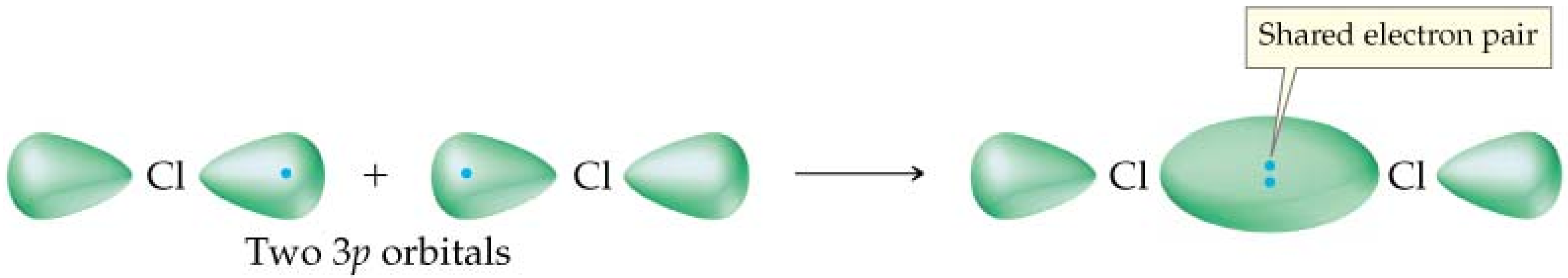


Fig 5.1 A covalent bond is the result of attractive and repulsive forces between atoms.

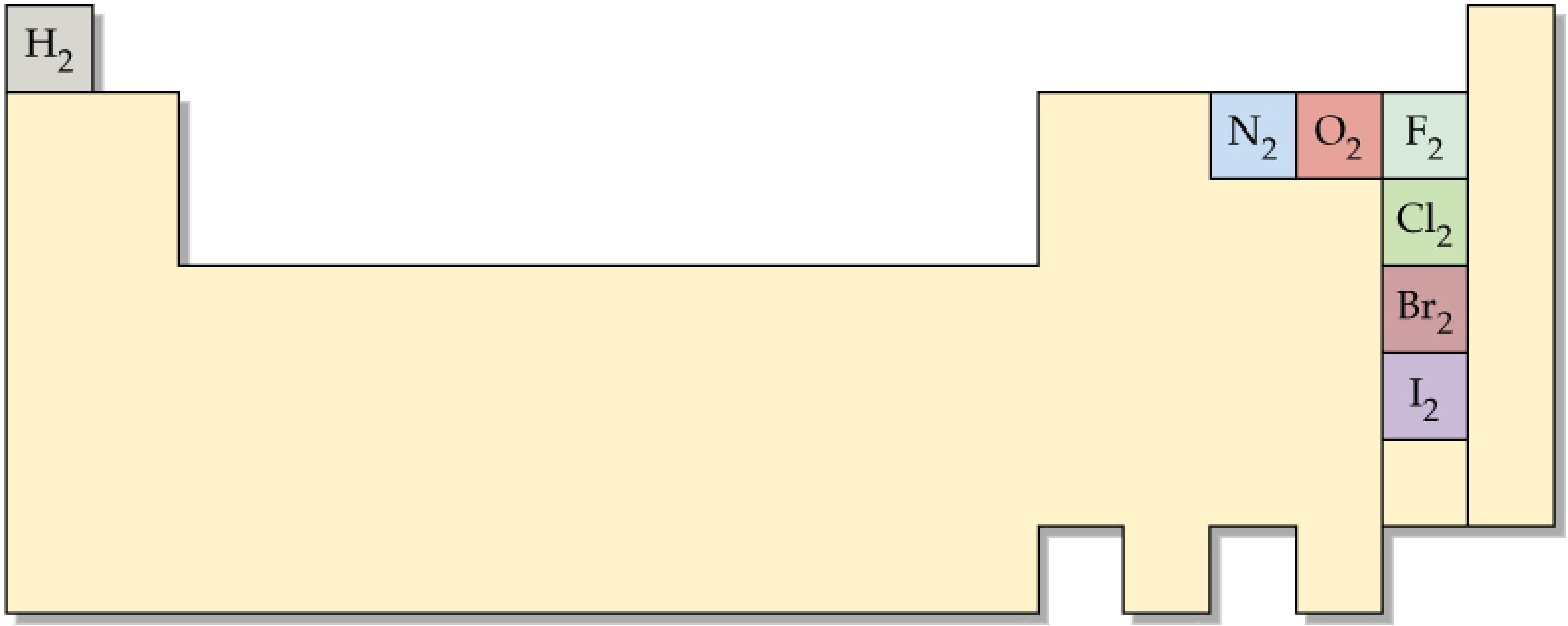
Spherical 1S orbital of two individual hydrogen atoms bends together and overlap to give an egg shaped region in the hydrogen molecule. The shared pair of electrons in a covalent bond is often represented as a line between atoms.



When two chlorine atoms approach each other, the unpaired 3*p* electrons are shared by both atoms in a covalent bond. Each chlorine atom in the Cl_2 molecule now have 6 electrons in its own valence shell and sharing two giving each valence shell octet.



In addition to H_2 and Cl_2, five other elements always exist as diatomic molecule.



5.2 Covalent Bonds and the Periodic Table

- Covalent bonds can form between unlike atoms as well as between like atoms, making possible a vast number of molecular compounds.
- Water molecule, H_2O, consists of two hydrogen atoms joined by covalent bonds to one oxygen atom.
- Ammonia molecule, NH_3, consists of three hydrogen atoms joined by covalent bonds to one nitrogen atom.
- In most covalent molecules, each atom shares enough electrons to achieve a noble gas configuration.

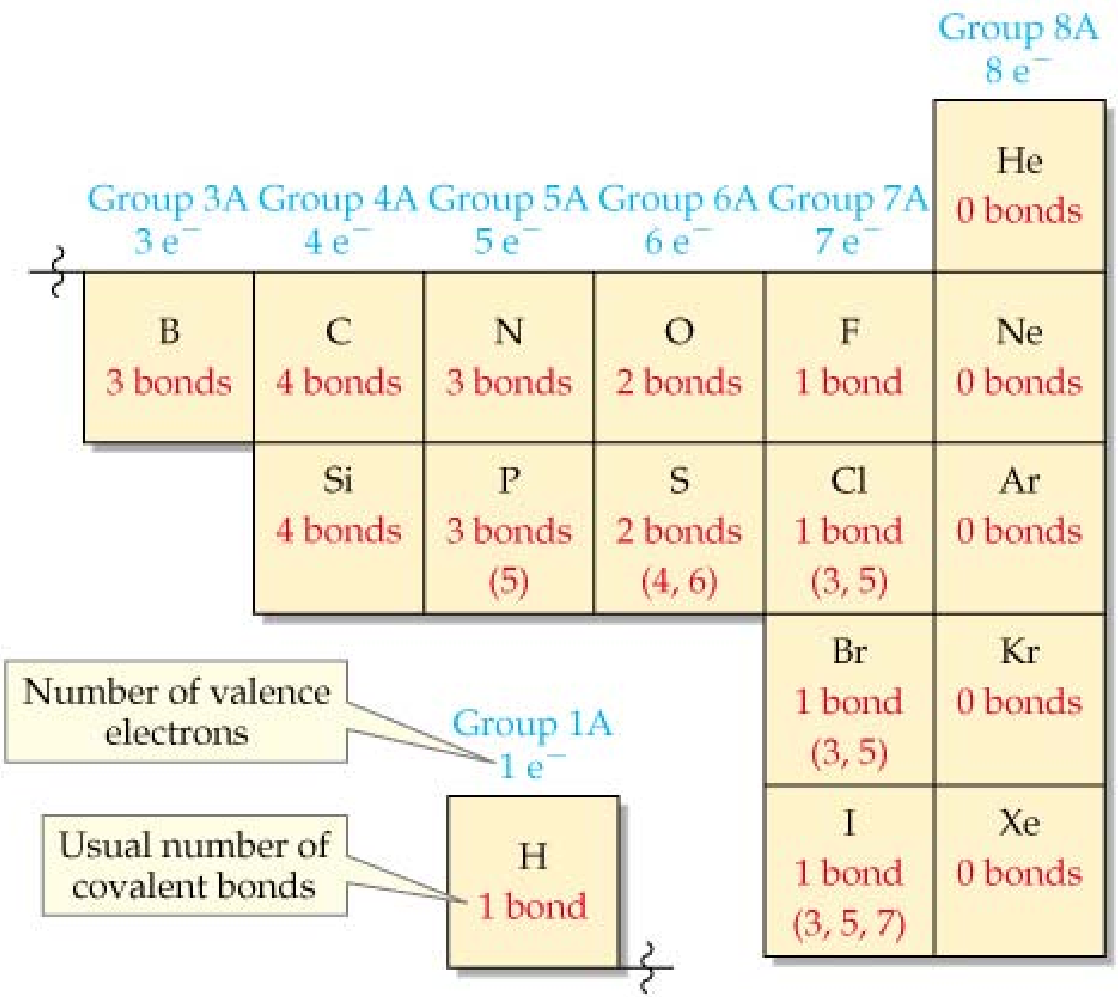
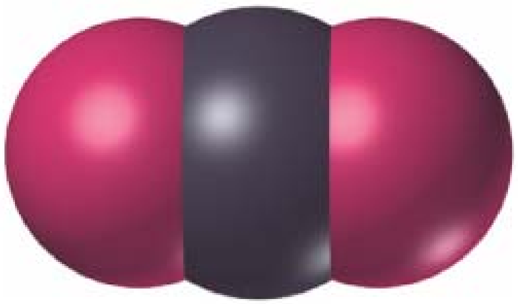


Fig 5.4 For P, S, Cl, and other elements in the third period and below, the number of covalent bonds may vary, as indicated by the numbers shown in parentheses.

5.3 Multiple Covalent Bonds

- ***Single bond***: A bond formed by sharing two electrons or one pair – represented by a single line between the atoms.
- ***Double bond***: A bond formed by sharing four electrons or two pairs – represented by two line between the atoms.

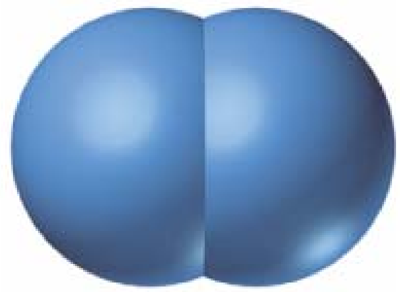
- ***Triple bond***: A bond formed by sharing six electrons or three pairs– represented by three line between the atoms.



Double bonds

$:\ddot{O}::C::\ddot{O}:$ *or* $:\ddot{O}=C=\ddot{O}:$

Carbon and two oxygen atoms have 8 electrons each.



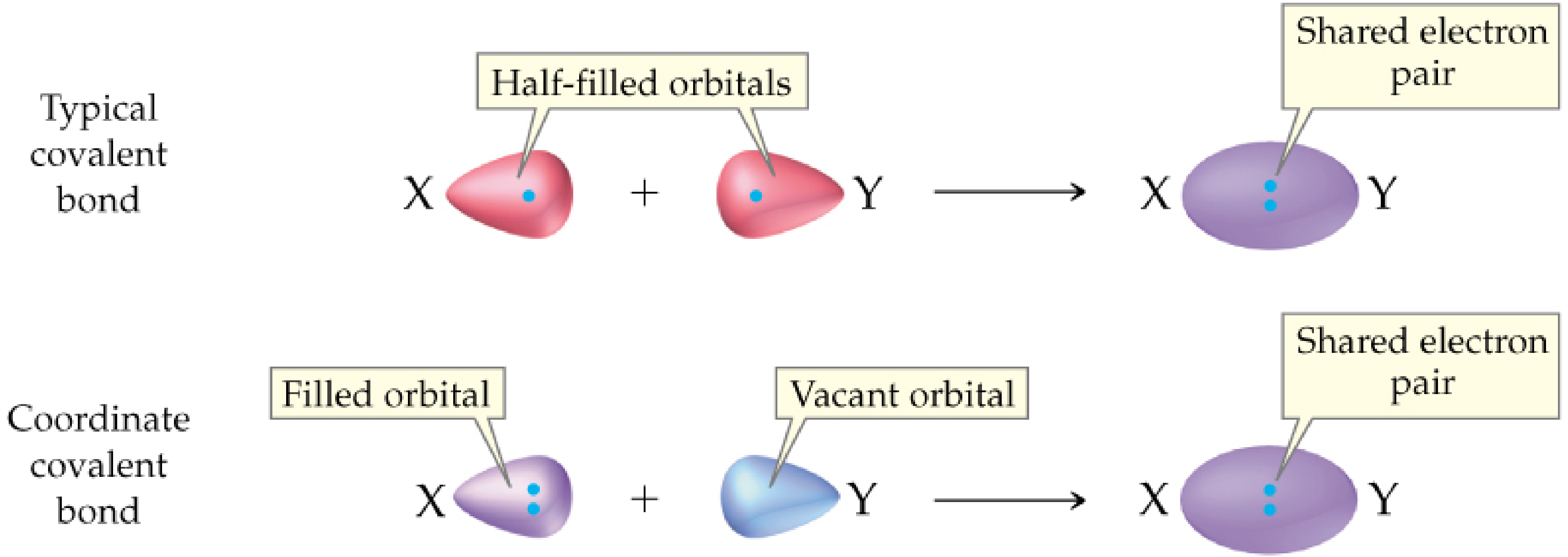
A triple bond

$:N\vdots\vdots N:$ *or* $:N\equiv N:$

Each nitrogen atom has 8 electrons.

5.4 Coordinate Covalent Bonds

Coordinate Covalent Bond: The covalent bond that forms when both electrons are donated by the same atom.



5.5 Molecular Formulas and Lewis Structures

Molecular Formula: A formula that shows the numbers and kind of atoms in one molecule of a compound.

Structural formula: A molecular representation that shows the connections among atoms by using lines to represent covalent bonds.

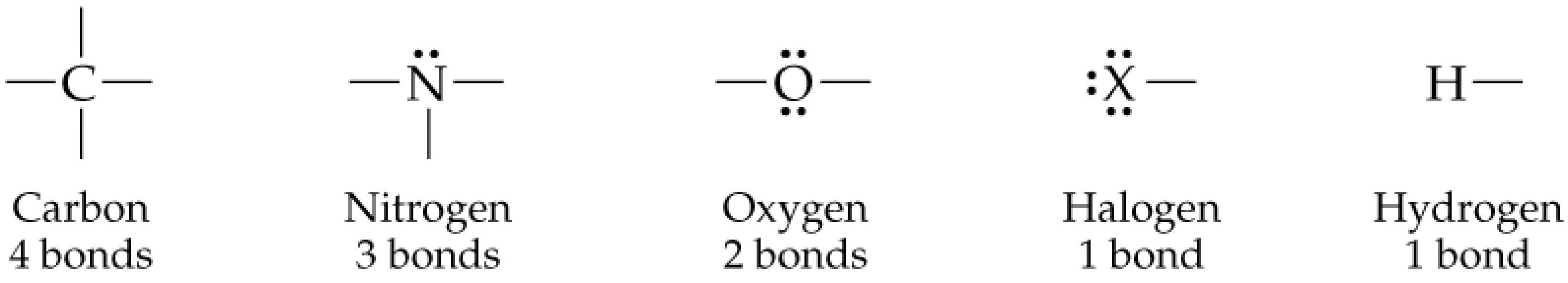
Lewis structure: A molecular representation that shows both the connections among atoms and the locations of lone-pair valence electrons.

Lone pair: A pair of electrons that is not used for bonding.

Condensed Structure: bonds are not specifically shown. For example, Ethane is written as CH_3CH_3 according to the condensed formula.

5.6 Drawing Lewis Structure

To draw Lewis structure, you need to know the connections among atoms. Also, common bonding patterns, shown below, for C, N, O, X, and H simplifies writing Lewis structure.



The following general method can be used to draw Lewis structure:

- Step 1: Find the total number of valence electrons of all atoms in the molecule or ion. For example, total number of valence electrons in PCl_3 is 26; phosphorus (Group 5A) has 5 valence electrons and each of 3 chlorine (Group 7A) has 7 valence electrons.
- Step 2: Draw a line between each pair of connected atoms to represent the two electrons in a covalent bond.

- Step 3: Add lone pairs so that each peripheral atom (except H) connected to the central atom gets an octet.
- Step 4: Place all remaining electrons on the central atom.
- Step 5: If the central atom does not yet have an octet after all the electrons have been assigned, take a lone pair from neighboring atom and form a multiple bond to the central atom.

5.7 Shape of Molecules

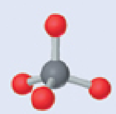
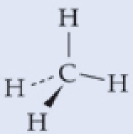
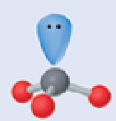
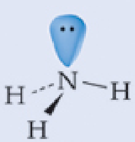
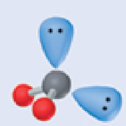
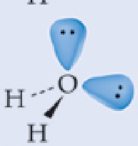
Molecular shapes can be predicted by noting how many bonds and electron pairs surround individual atoms and applying what is called the ***valence-shell electron-pair repulsion*** *(VSEPR)* model. The basic idea of VSEPR model is that the negatively charged clouds of electron in bonds and lone pair repel each other therefore tends to keep apart as far as possible causing molecules to assume specific shape.

There are three step to applying the VSEPR model:

- Step 1: Draw a Lewis structure of the molecule, and identify the atom whose geometry is of interest.
- Step 2: Count the number of electron charge clouds surrounding the atom of interest.
- Step 3: Predict molecular shape by assuming that the charge clouds orient in space so that they are as far away from one another as possible.

The shape depends on the number of charged clouds surrounding the atom as summarized in Table 5.1

TABLE 5.1 Molecular Geometry Around Atoms with 2, 3, and 4 Charge Clouds

Number of Bonds	Number of Lone Pairs	Number of Charge Clouds	Molecular Geometry	Example
2	0	2	 Linear	O=C=O
3	0	3	 Planar triangular	
2	1		 Bent	
4	0	4	 Tetrahedral	
3	1		 Pyramidal	
2	2		 Bent	

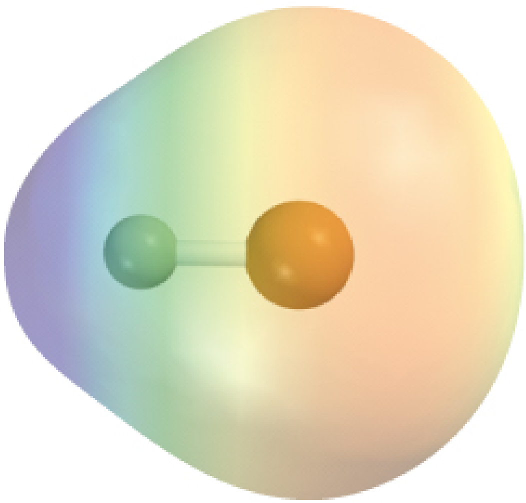
5.8 Polar Covalent Bonds and Electronegativity

- Electrons in a covalent bond occupy the region between the bonded atoms.
- If the atoms are identical, as in H_2 and Cl_2, electrons are attracted equally to both atoms and are shared equally.
- If the atoms are not identical, however, as in HCl, the bonding electrons may be attracted more strongly by one atom than by the other and may thus are shared unequally. Such bonds are known as *polar covalent bonds*.

This end of the molecule is electron-poor and has a partial positive charge ($\delta+$).

This end of the molecule is electron-rich and has a partial negative charge ($\delta-$).

$$\overset{\delta+}{\mathrm{H}}-\overset{\delta-}{\mathrm{Cl}}$$



- In HCl, electron spend more time near the chlorine than the hydrogen. Although the molecule is overall neutral, the chlorine is more negative than the hydrogen, resulting in partial charges on the atoms.
- Partial charges are represented by placing δ- on the more negative atom and δ+ on the more positive atom.
- Ability of an atom to attract electrons is called the atom's *electronegativity*.
- Fluorine, the most electronegative element, assigned a value of 4, and less electronegative atoms assigned lower values, as shown in fig 5.7.

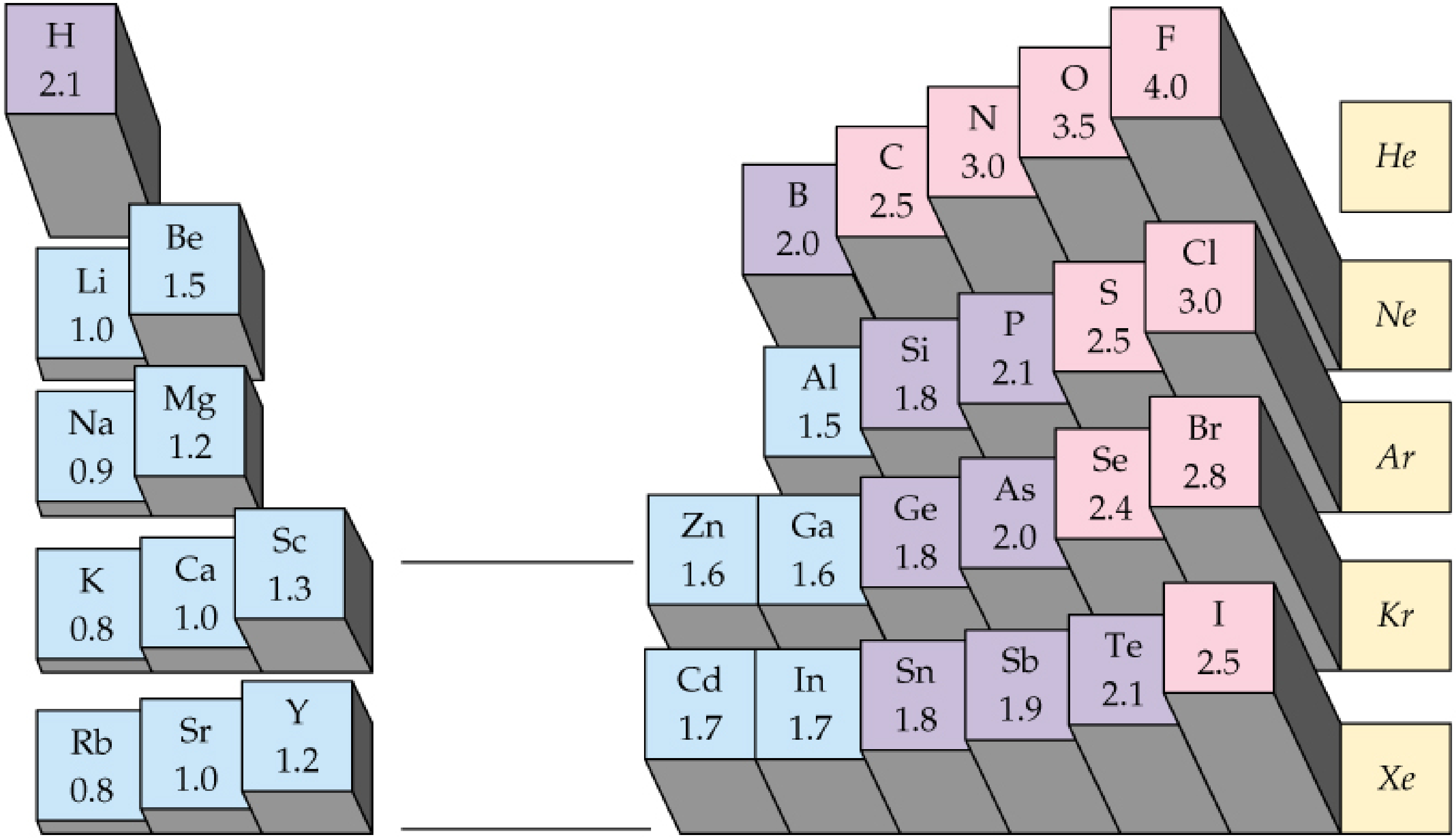
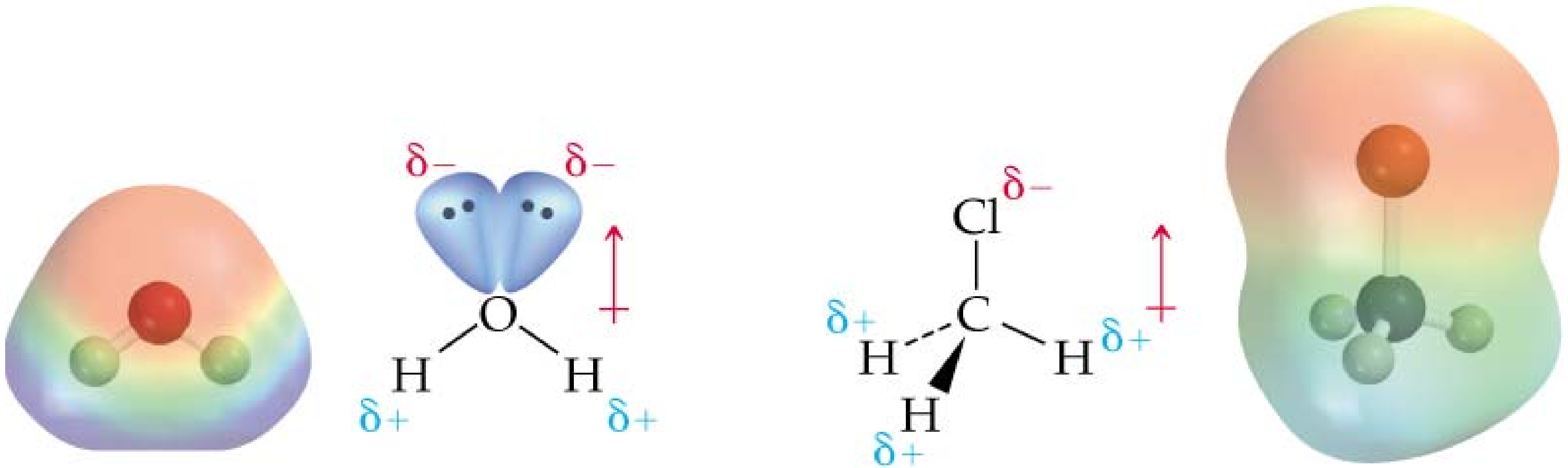


Fig 5.7 Electronegativities and the periodic table

5.9 Polar Molecules

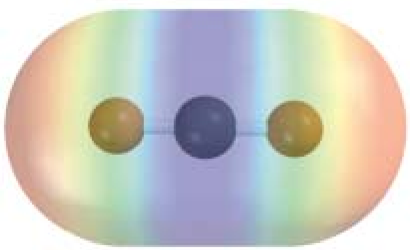
- Entire molecule can be polar if electrons are attracted more strongly to one part of the molecule than to another.
- Molecule's polarity is due to the sum of all individual bond polarities and lone-pair contribution in the molecule.

- Molecular polarity is represented by an arrow pointing at the negative end and is crossed at the positive end to resemble a positive sign.

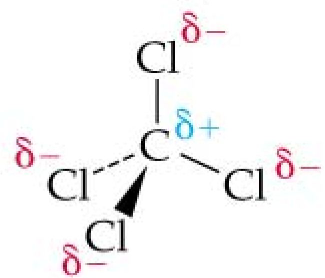


Water, H_2O Chloromethane, CH_3Cl

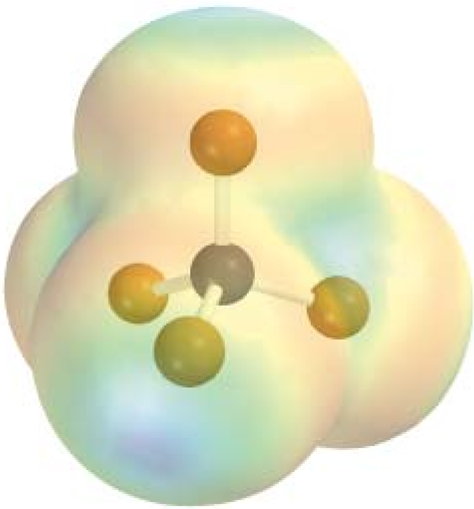
- Molecular polarity depends on the shape of the molecule as well as the presence of polar covalent bonds and lone-pairs.



Zero net polarity



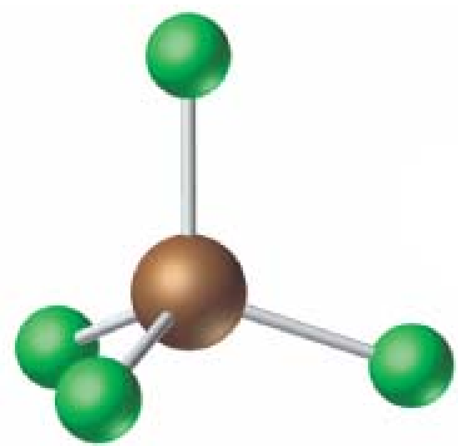
Zero net polarity



5.10 Naming Binary Molecular Compounds

- When two different elements combines together they form *binary compound*.
- The formulas of binary compounds are usually written with the less electronegative element first. Thus, metals are always written before non-metals. Prefix such as mono, di, tri, tetra etc, are used to indicate number of atoms of each element.

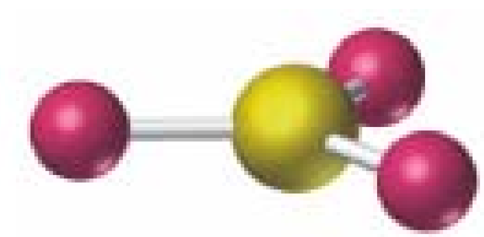
A few examples of binary compounds are given below:



$TiCl_4$

BCl_3

NO_2



SO_3

The following two steps guide is helpful in naming binary compounds:

- Step 1: Name the first element in the formula, using a prefix if needed to indicate the number of atoms.
- Step 2: Name the second element in the formula, using an –ide ending as for anions, along with a prefix if needed to indicate the number of atoms.

5.11 Characteristic of Molecular Compounds

Molecules are neutral as a result there is no strong electrical attractions between the molecules to hold them together. However, there are several weaker forces exist between molecules, known as intermolecular forces .

- When intermolecular forces are very weak, molecules are weakly attracted to one another and that the substance is gas at ordinary temperature.

- If the intermolecular forces are somewhat stronger, the molecules are pulled together into a liquid.
- If the forces are stronger, the substance becomes a molecular solid.
- Melting points and boiling points of molecular solids are lower than those of ionic solids.
- Most molecular compounds are insoluble in water.
- Molecular compounds do not conduct electricity when melted because they have no charged particles.