Chemical Bonding

Chemistry

In this lesson, we will discuss the following:

- **Electron-Dot Structures**
- **Ions**
- **Ionic Bonds**
- **Covalent Bonds**

Electron-Dot Structures

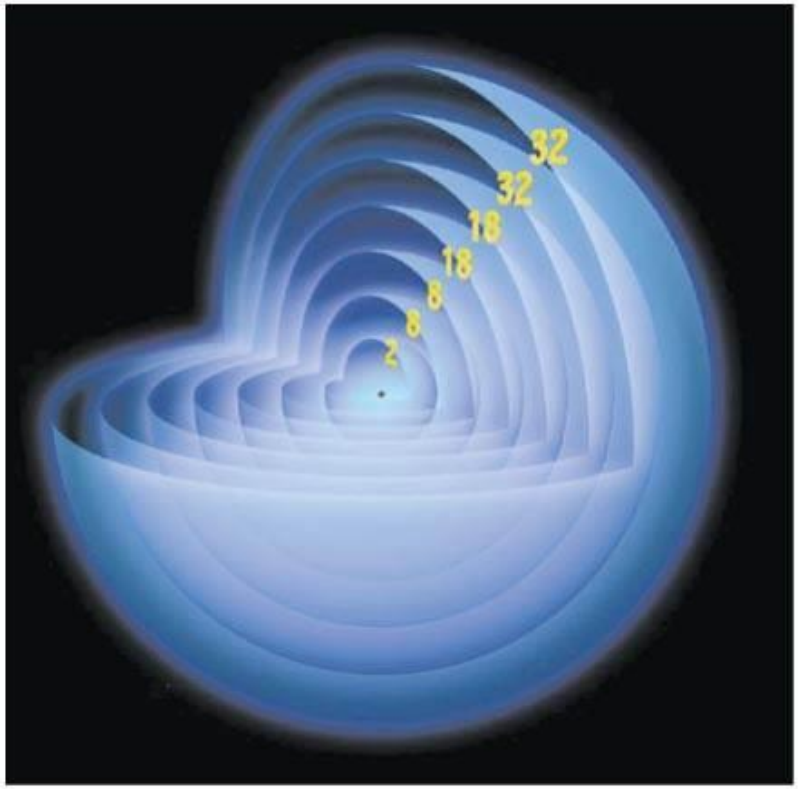
- Atoms bond by using electrons.
- In order to understand chemical bonding, we then need to understand a little bit about how electrons are arranged in an atom.

Electron-Dot Structures

- The electrons in an atom behave as if they are arranged in concentric shells, called orbits.

Electron-Dot Structures

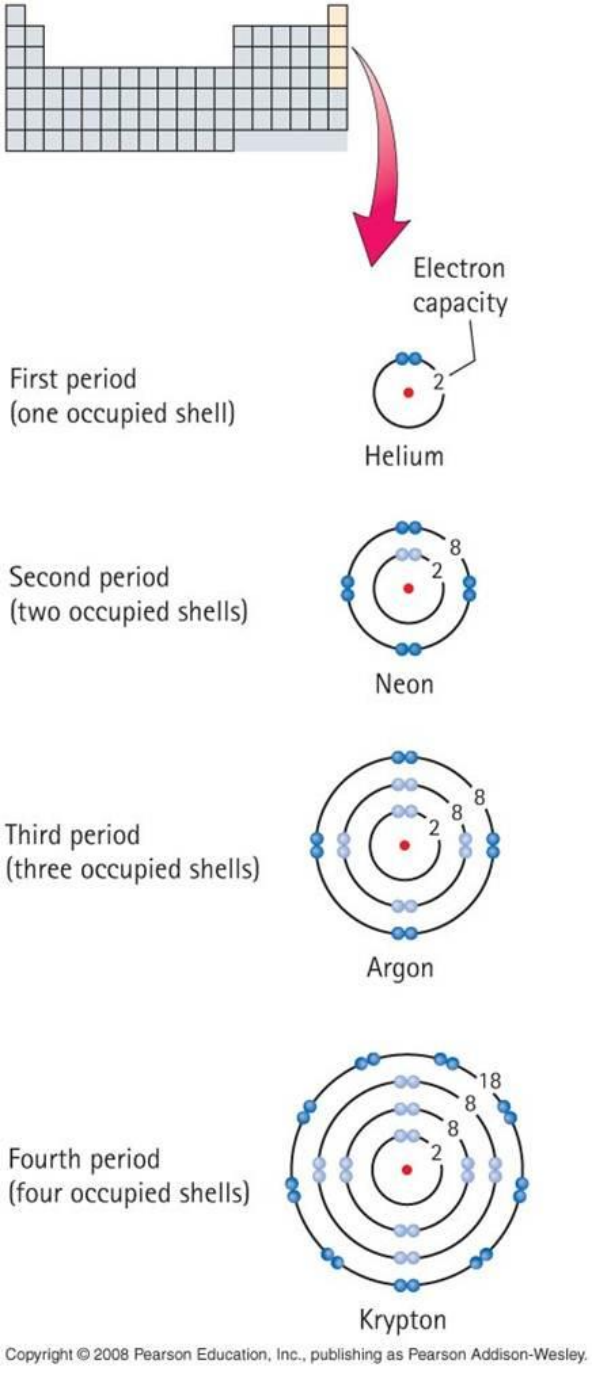
- The different electron shells can hold various maximum amounts of electrons.
  - The innermost electron shell can hold a maximum of 2 electrons.
  - The next two shells can hold a maximum of 8 electrons.
  - The next two shells can hold a maximum of 18 electrons.
  - The next two shells can hold a maximum of 32 electrons.



Electron-Dot Structures

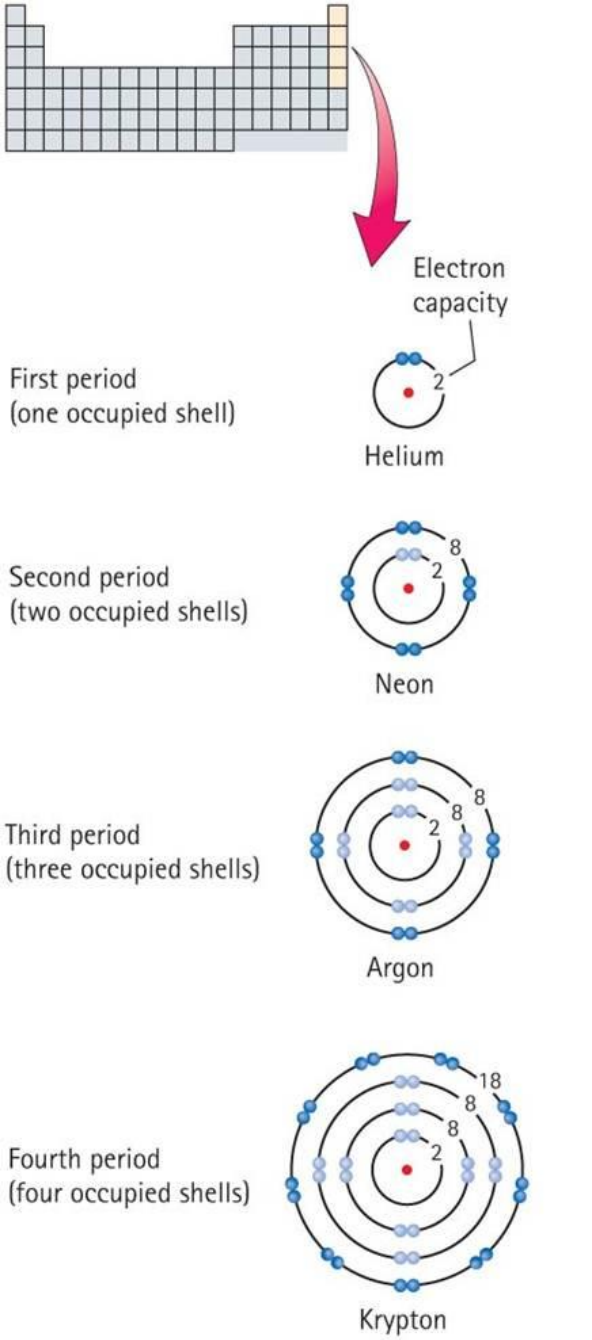
- If we draw the shells in two dimensions, the atoms look like the picture at the right.

The atoms shown at right all come from the same group (or family).



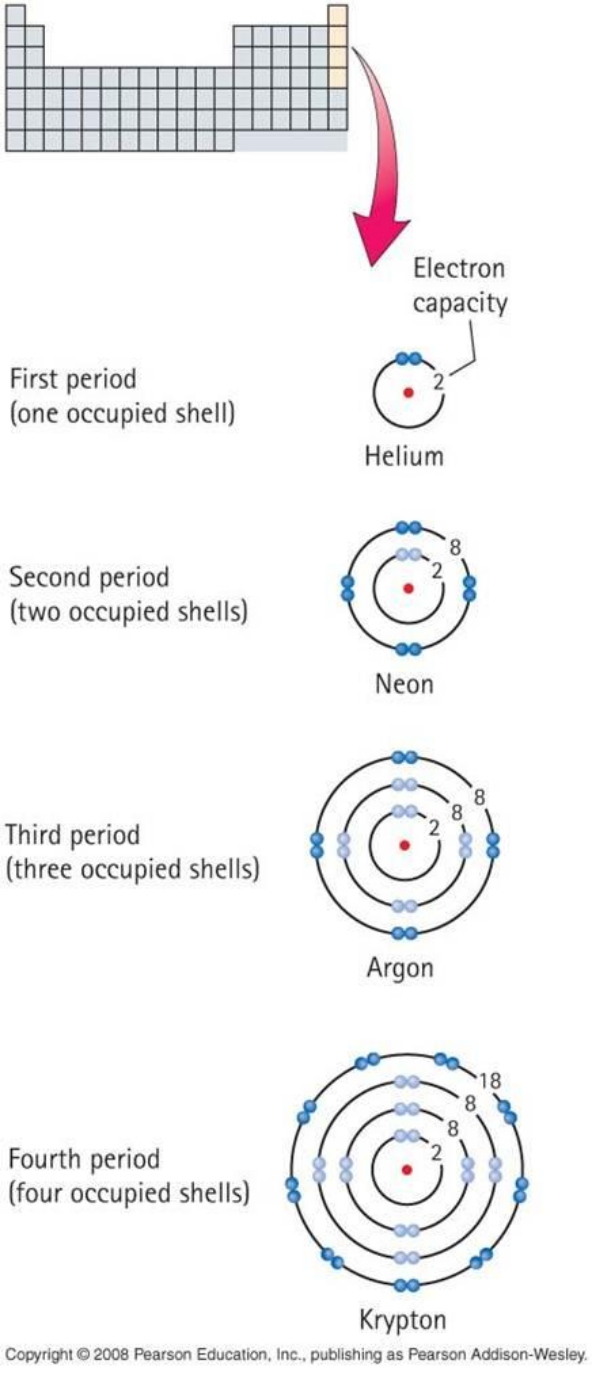
Electron-Dot Structures

- Only the electrons from the outermost shell of an atom can participate in chemical bonding.
- These electrons are called **VALENCE ELECTRONS**.



Electron-Dot Structures

- While heavier elements have more than 8 valence electrons, only a maximum of 8 of the outermost electrons are able to participate in bonding.
- Thus, we say that heavy elements can only have a maximum of 8 outermost electrons.



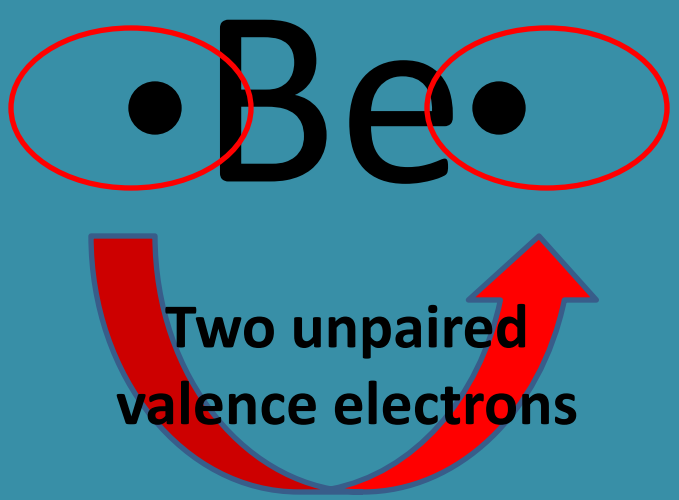
Electron-Dot Structures

- Electron-dot structures are important because they let us know which electrons can participate in chemical bonding.
- Electron-dot structure show only an atoms valence electrons, which are the only electrons which participate in bonding.
- Electron-dot structures also show valence electrons as either in pairs or not in pairs. Only **unpaired** valence electrons can participate in **covalent** bonding.

Electron-Dot Structures

- To draw the electron-dot structure for an element, do the following:
  - Write down the elemental symbol. The symbol has 4 sides.
  - Draw the outermost electrons around the symbol. Put one electron on each of the 4 sides before pairing electrons up on any side. The exception is Helium. Its two electrons should be drawn as a pair.

Example:



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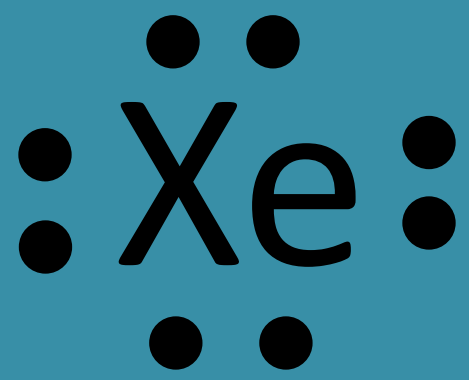
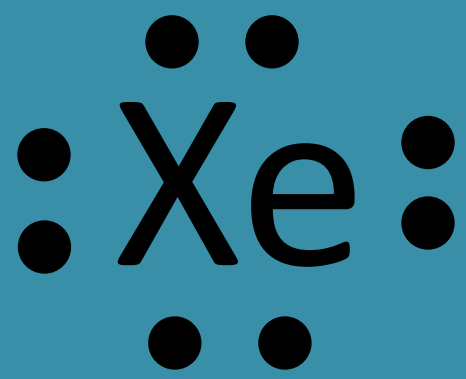


Electron-Dot Structures

No unpaired valence electrons: no covalent bonds. In addition, the valence shell for xenon is full (has eight electrons), and so xenon will not form ionic bonds, either.

Example:



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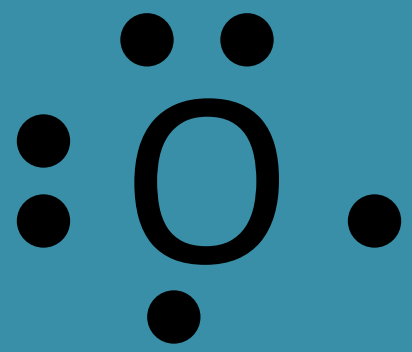
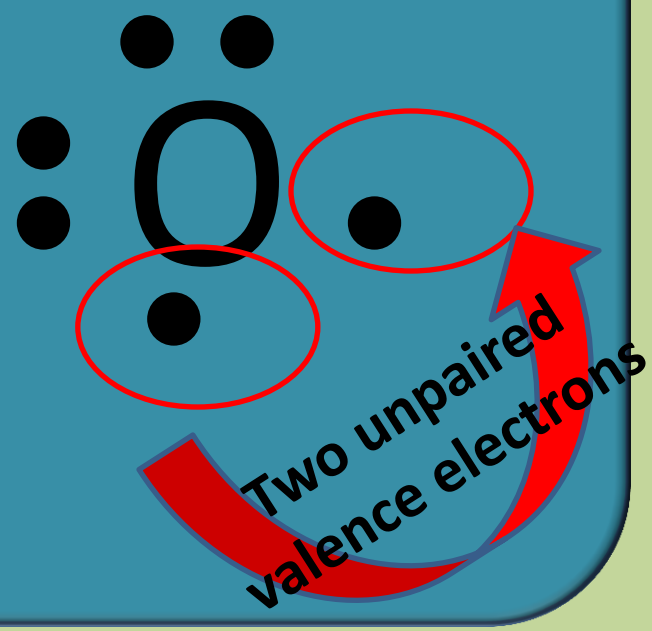


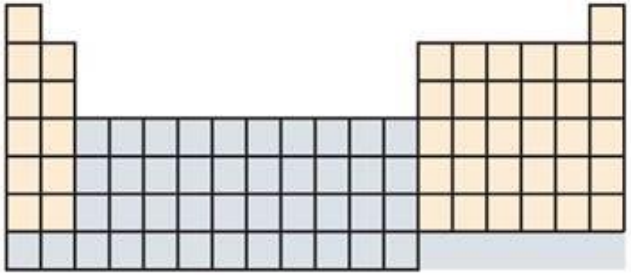
Electron-Dot Structures

Two unpaired valence electrons: oxygen can make at most two covalent bonds. Because its valence shell is not full (it does not have eight electrons), it will form an ionic bond.

Example:



Electron-Dot Structures



1	2	13	14	15	16	17	18
H·							He:
Li·	·Be·	·B·	·C·	·N·	:O·	:F·	:Ne:
Na·	·Mg·	·Al·	·Si·	·P·	:S·	:Cl·	:Ar:
K·	·Ca·	·Ga·	·Ge·	·As·	:Se·	:Br·	:Kr:
Rb·	·Sr·	·In·	·Sn·	·Sb·	:Te·	:I·	:Xe:
Cs·	·Ba·	·Tl·	·Pb·	·Bi·	:Po·	:At·	:Rn:

Ions

- An ion is defined as an atom which has lost or gained one or more electrons.

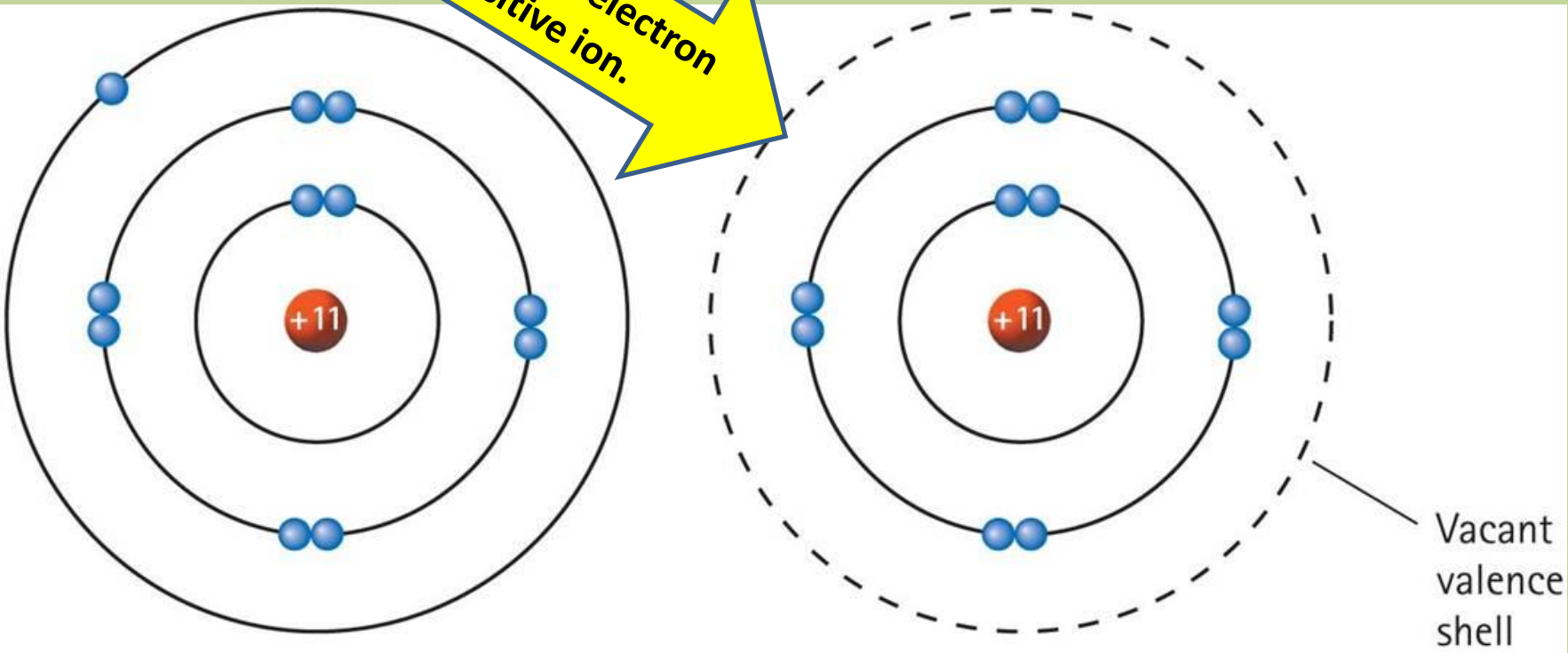
Click HERE to learn a little bit about ionic bonding.

Ions

- An ion is defined as an atom which has lost or gained one or more electrons.
- If an atom has lost one or more electrons, it is positively-charged and is called a positive ion, or a cation.

METALS tend to lose electrons and become POSITIVE IONS.

Sodium tends to lose an electron to become a positive ion.



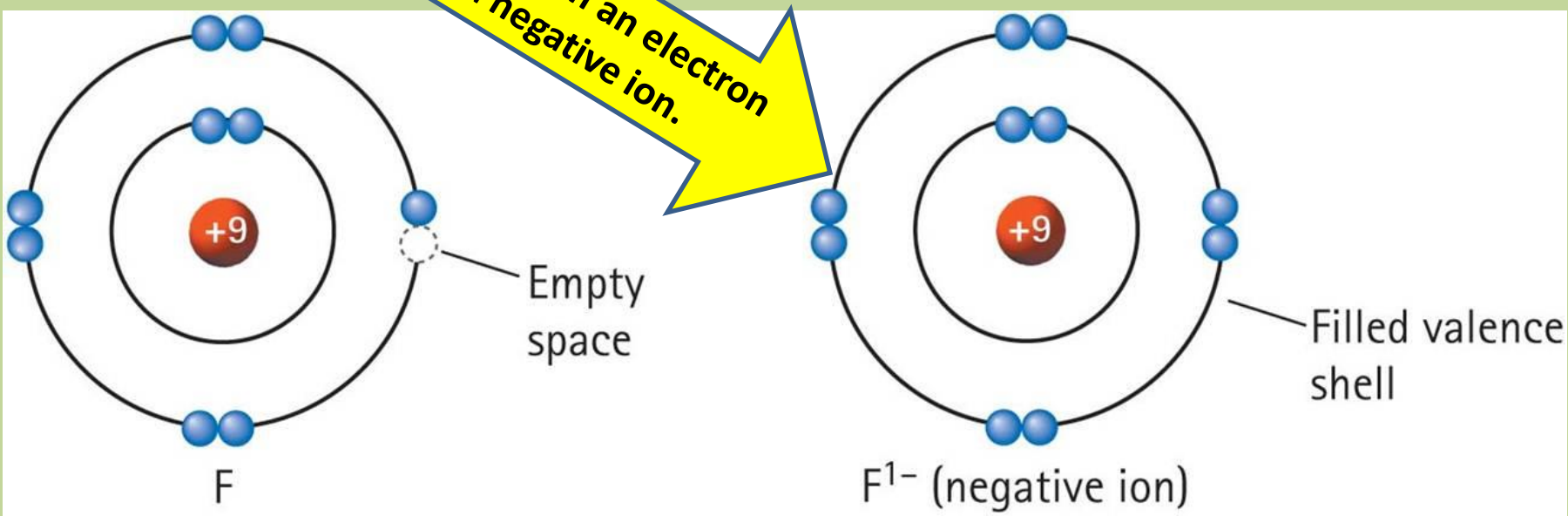
Na	Na^{1+} (positive ion)
11 protons	11 protons
11 electrons	10 electrons
0 net charge	+1 net charge

Ions

- An ion is defined as an atom which has lost or gained one or more electrons.
- If an atom has gained one or more electrons, it is negatively-charged and is called a negative ion, or an anion.

NONMETALS tend to gain electrons and become NEGATIVE IONS.

Fluorine tends to gain an electron to become a negative ion.



F

9 protons
9 electrons
0 net charge

F^{1-} (negative ion)

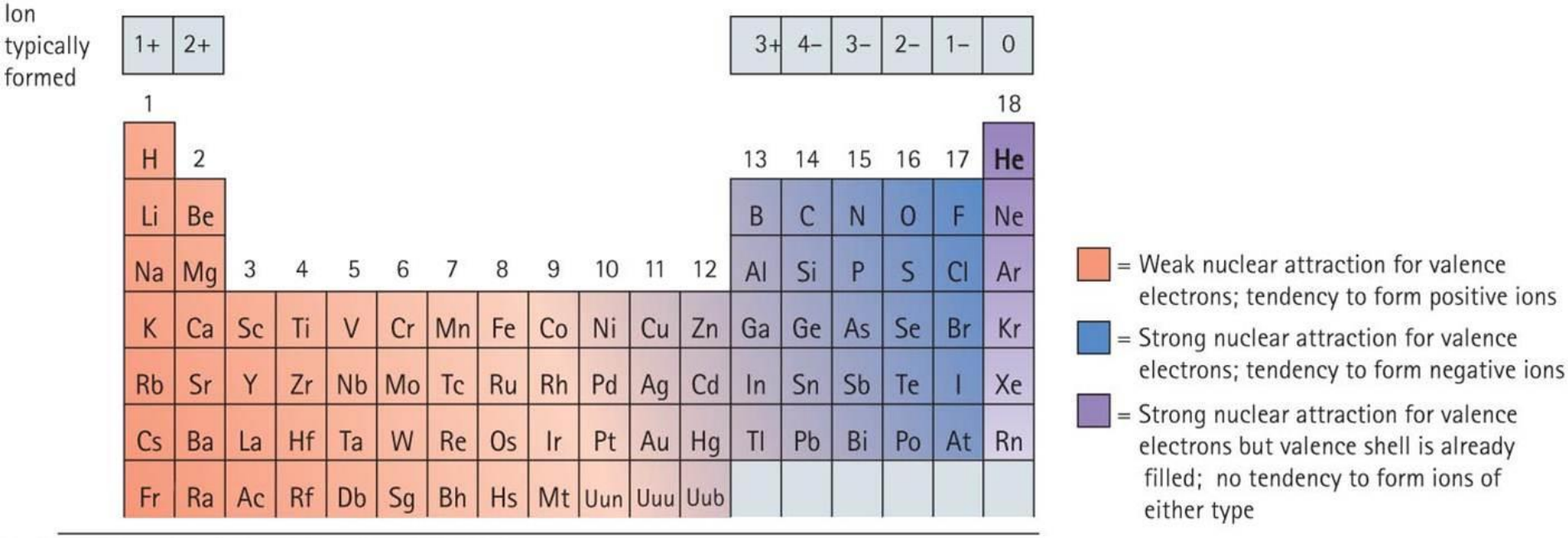
9 protons
10 electrons
−1 net charge

Ions

- An ion is defined as an atom which has lost or gained one or more electrons.

In general, when becoming ions or when bonding, atoms like to either completely fill their outermost shells or completely empty their outermost shells, whichever is easier.

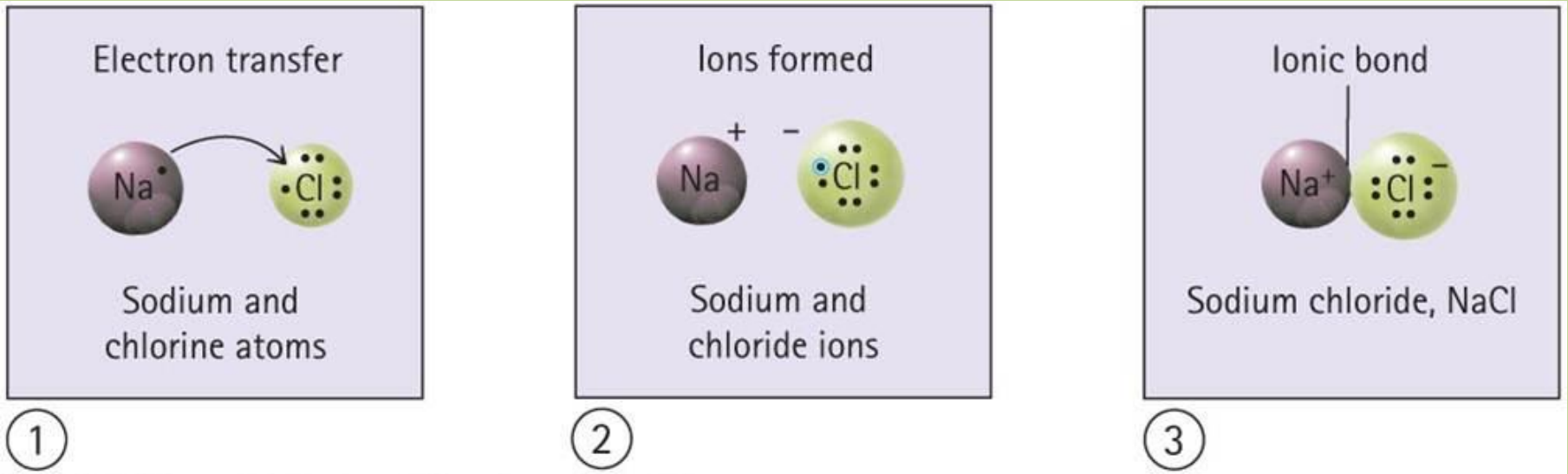
Ions



We say that atoms want to achieve a stable OCTET. In other words, atoms want 8 outer electrons. Helium, hydrogen, lithium, beryllium, and boron are exceptions; they want only 2 outer electrons.

Ionic Bonds

- An ionic bond is just the electric force of attraction between oppositely-charged ions.



Ionic Bonds

- An ionic bond is just the electric force of attraction between oppositely-charged ions.
- Metals tend to become positively-charged ions, and nonmetals tend to become negatively-charged ions. In other words, metals and nonmetals become oppositely-charged ions.
- In other words, an ionic bond forms between a metal and a nonmetal.

Ionic Bonds

- To write the formula for an ionic compound, do the following:
  - Write down the metal ion with its charge first.
  - Write down the nonmetal ion with its charge next.
  - Crisscross charges to get subscripts. Lose the charge signs.
  - Reduce subscripts if possible.

Example: Write the formula for beryllium oxide.

$$Be^{+2}\ O^{-2}$$

$$1 \quad 1$$

Covalent Bonds

- A covalent bond is defined as the attraction of two atoms based on a pair of shared electrons.
- One covalent bond consists of TWO electrons shared by two atoms.

Click HERE to learn a little bit about covalent bonding.

Covalent Bonds

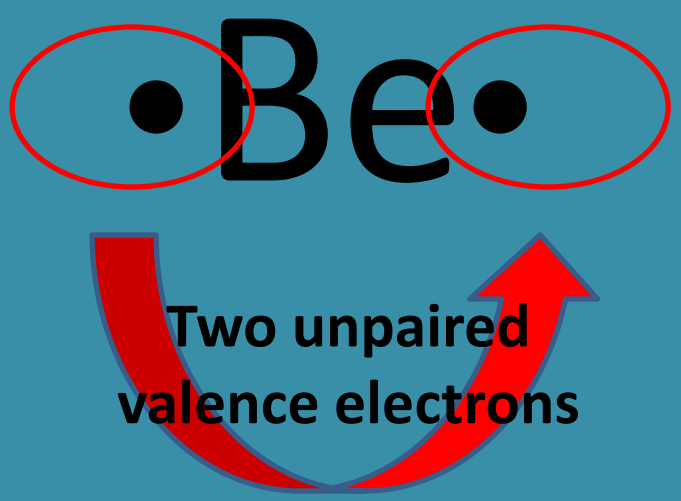
- Covalent bonds occur between nonmetals.
- Metals do not participate in covalent bonding because metals simply give up their electrons. Covalent bonds require SHARING, not giving of electrons.

Click HERE to watch a short video about covalent bonding.

Covalent Bonds

- The number of covalent bonds an atom is able to make is equal to the number of unpaired electrons in its electron-dot structure.

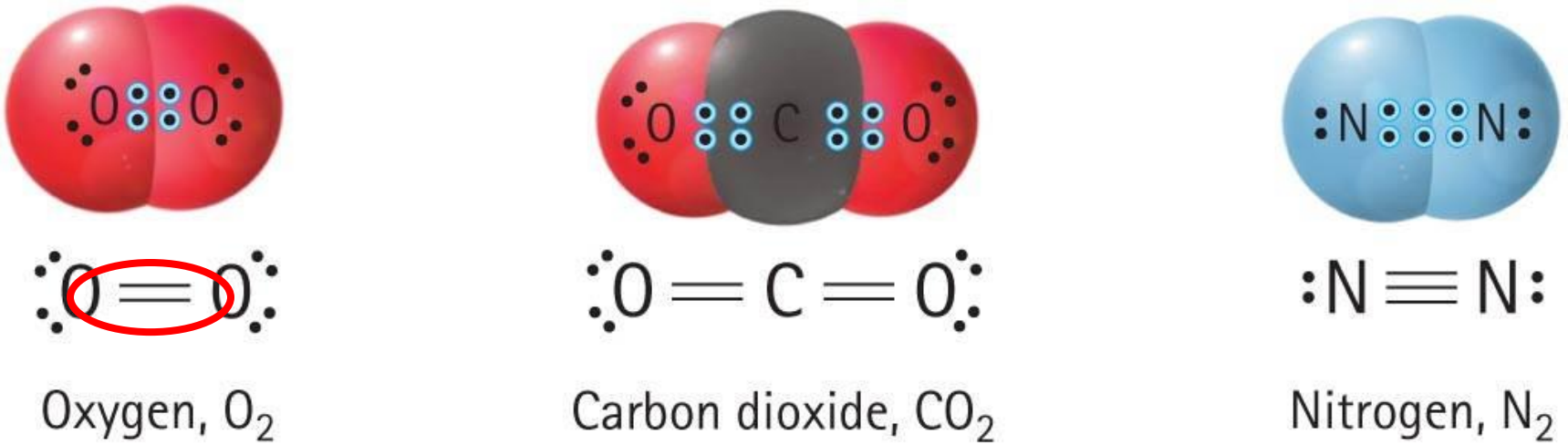
Example:



Covalent Bonds

- It is possible for there to be multiple covalent bonds between two atoms.

A single covalent bond is often represented with a single line.

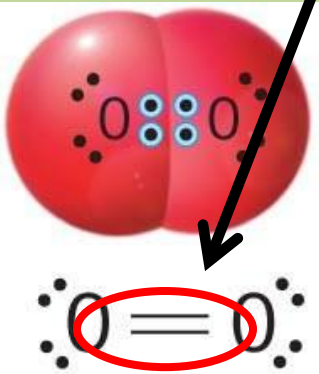


valent Bonds

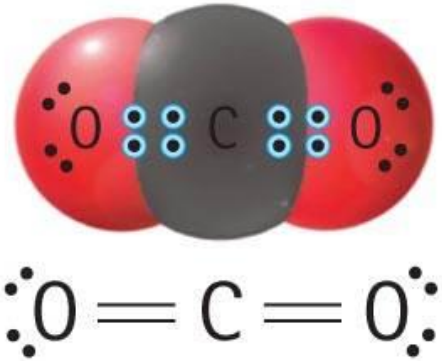
- It is p here to be multiple covalent bond vo atoms.

Multiple covalent bonds are represented with multiple lines. For instance, two lines means two covalent bonds.

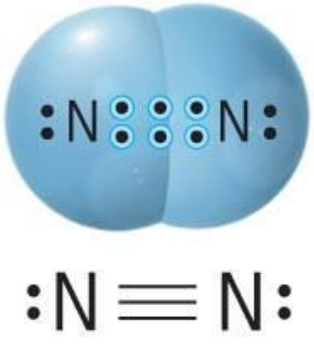
A single covalent bond is often represented with a single line.



Oxygen, O_2



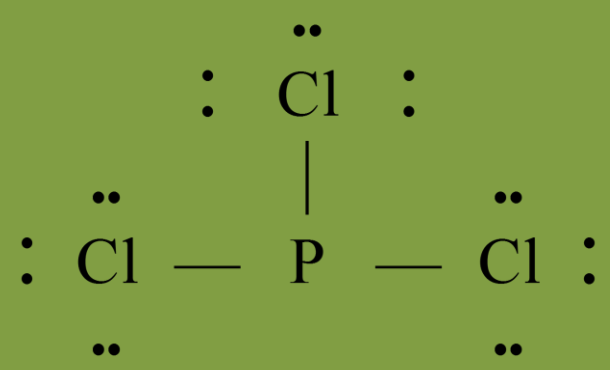
Carbon dioxide, CO_2



Nitrogen, N_2

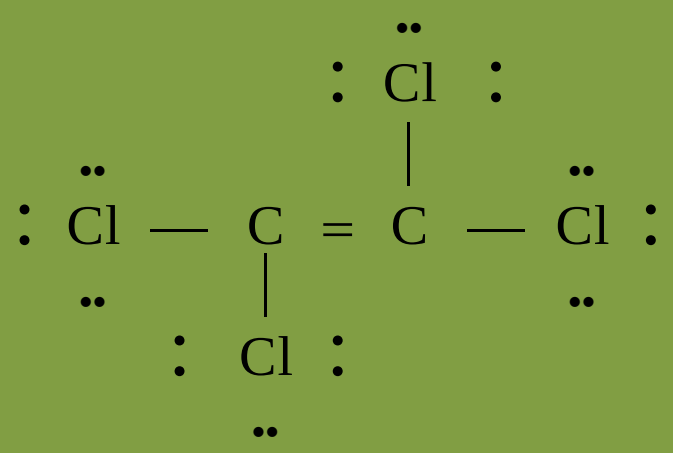
Writing Lewis Structures for Covalent Molecules

1. Attach the atoms together in a skeletal structure.
   - Most metallic element is generally central.
     - In PCl_3, the P is central because it is farther left on the periodic table and therefore more metallic.



Writing Lewis Structures for Covalent Molecules

1. Attach the atoms together in a skeletal structure.
   - Halogens and hydrogen are generally terminal.
     - In C_2Cl_4, the Cs are attached together in the center and the Cls are surrounding them.
   - Many molecules tend to be symmetrical.
     - Though there are many exceptions to this, chemical formulas are often written to indicate the order of atom attachment.
     - In C_2Cl_4, there are two Cls on each C.



Writing Lewis Structures for Covalent Molecules, Continued

2. Calculate the total number of valence electrons available for bonding.
   - Use group number of periodic table to find number of valence electrons for each atom.

Writing Lewis Structures for Covalent Molecules, Continued

3. Attach atoms with pairs of electrons and subtract electrons used from total.
   - Bonding electrons.
4. Add remaining electrons in pairs to complete the octets of all the atoms.
   - Remember H only wants 2 electrons.
   - Don’t forget to keep subtracting from the total.
   - Complete octets on the terminal atoms first, then work toward central atoms.

Writing Lewis Structures for Covalent Molecules, Continued

5. If there are not enough electrons to complete the octet of the central atom, bring pairs of electrons from an attached atom in to share with the central atom until it has an octet.
   - Try to follow common bonding patterns.

Example HNO_3

1. Write skeletal structure. H is on the outside, and N is central.

O

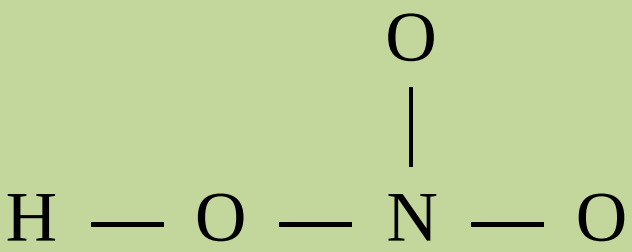
H O N O

2. Count valence electrons.

N = 5
H = 1
O_3 = 3·6 = 18
Total = 24 e^-

Example HNO_3 , Continued

3. Attach atoms with pairs of electrons and subtract from the total.

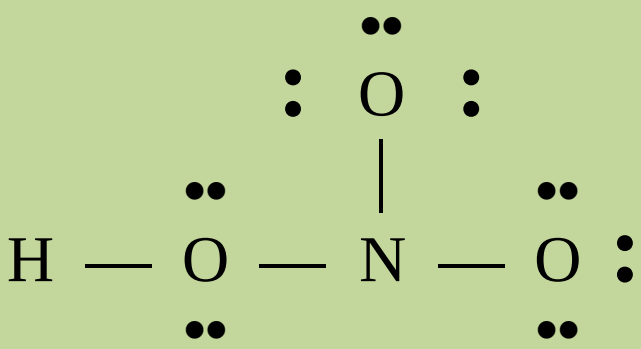


N = 5
H = 1
$O_3 = 3 \cdot 6 = 18$
Total = 24 e^-

	Electrons
Start	24
Used	8
Left	16

Example HNO_3 , Continued

4. Complete octets, outside-in.
   - H is already complete with 2.
     - 1 bond.
   - Keep going until all atoms have an octet or you run out of electrons.



		Electrons		Electrons
N = 5				
H = 1	Start	24	Start	16
O_3 = 3·6 = 18	Used	8	Used	16
Total = 24 e^-	Left	16	Left	0

Example HNO_3 , Continued

5. If central atom does not have octet, bring in electron pairs from outside atoms to share.
   - Follow common bonding patterns if possible.