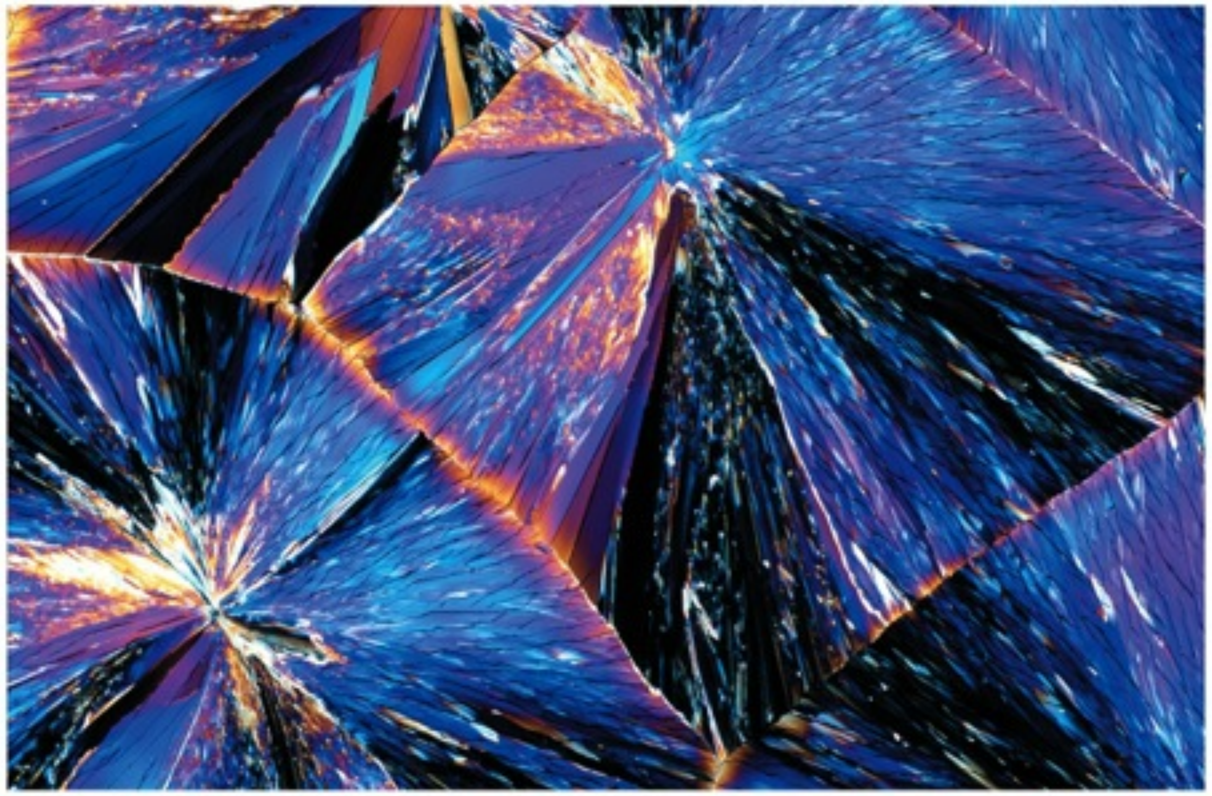
11 Chemical Bonds: The Formation of Compounds from Atoms



Pasieka/Photo Researchers, Inc.

Atoms in Vitamin C (ascorbic acid) bond in a specific orientation which defines the shape of the molecule. The molecules pack in a crystal, photographed above in a polarized micrograph (magnification 200x).

Chapter Outline

Periodic Trends in Atomic Properties

Periodic trends permit us to predict chemical properties and reactivity of the elements.

Metal properties: usually lustrous, malleable and good heat/electrical conductors.

Tend to lose electrons to form cations.

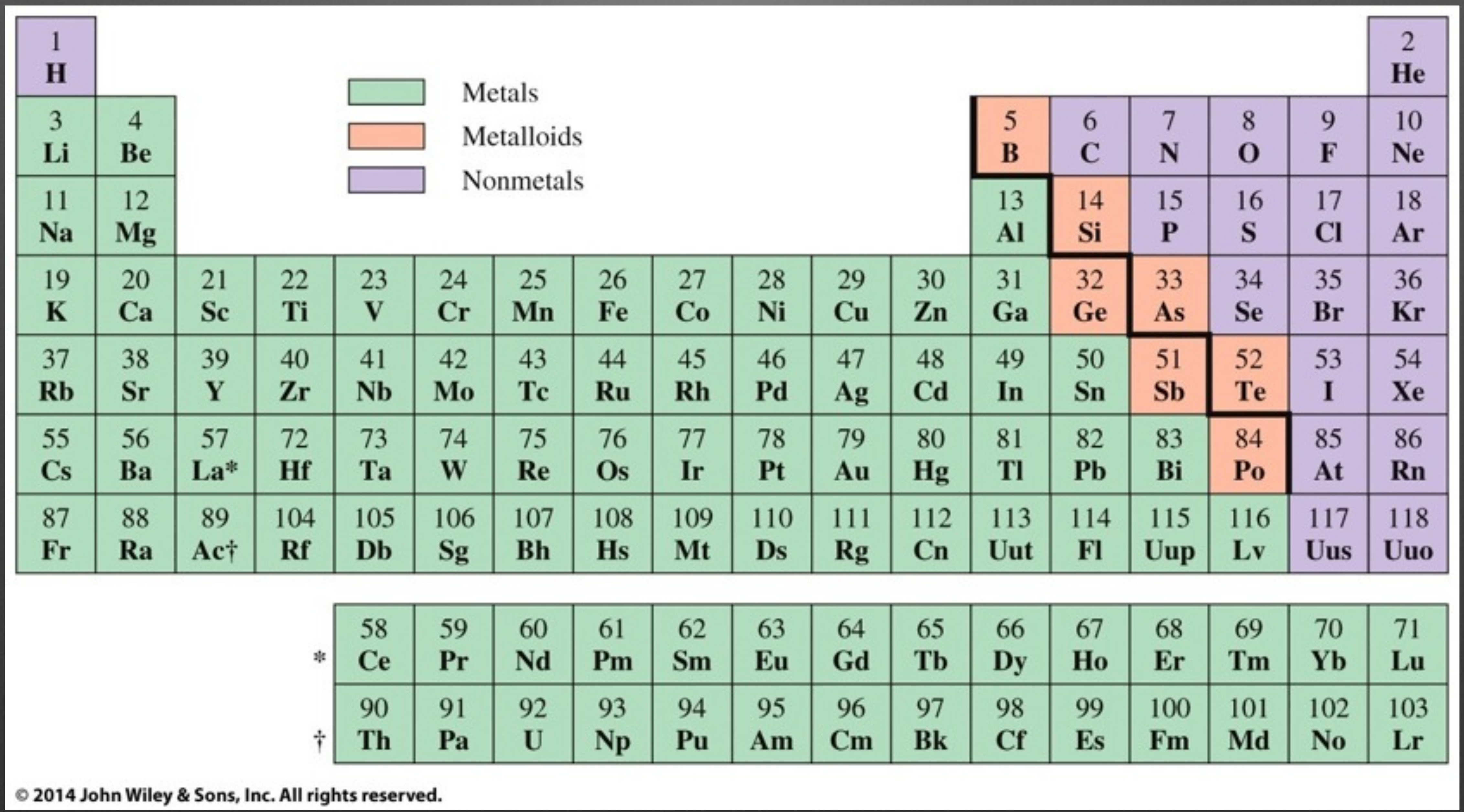
Nonmetal properties: usually not lustrous, brittle and poor heat/electrical conductors.

Tend to gain electrons to form anions.

Metalloid properties: can have some properties of either metals or nonmetals.

Periodic Trends in Atomic Properties

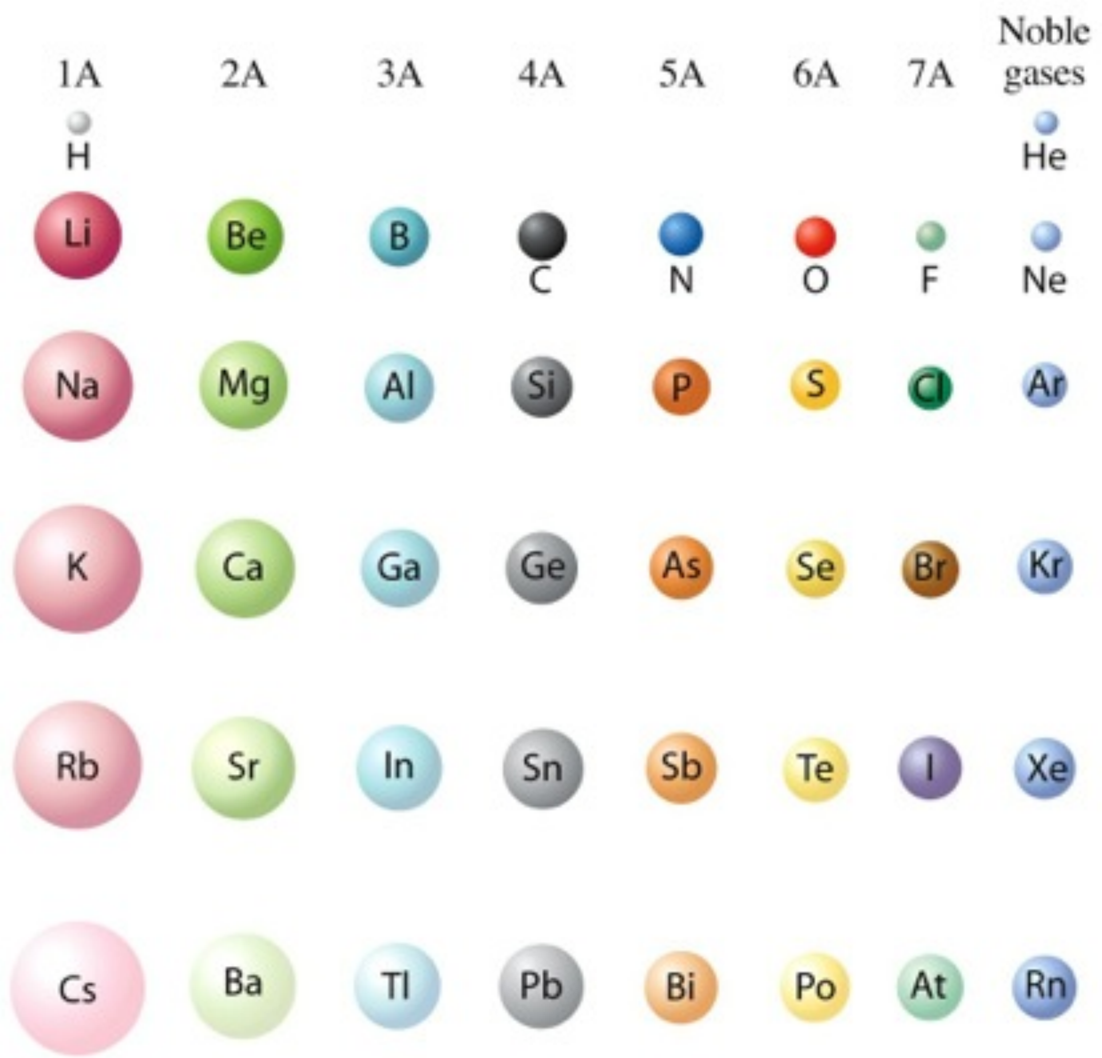
Classification of Metals, Nonmetals, and Metalloids



Hydrogen is an exception – properties more like a nonmetal.

Periodic Trends in Atomic Properties

Atomic Radius: increases down a group and decreases left to right across a period.



Atomic Radii for the Main Group Elements

Periodic Trends in Atomic Properties

Atomic Radius

General trend: increases down a group and decreases left to right across a period.

Down a Group:
Additional n quantum levels are added; electrons are farther from the nucleus, so size increases.

Across a Period:
Left to right, n remains constant but atomic number increases; increased nuclear charge creates greater interaction with electrons, resulting in a decrease in radii.

Periodic Trends in Atomic Properties

Ionization Energy: Energy required to remove an electron from an atom.

$$Na + \text{ionization energy} \longrightarrow Na^+ + e^-$$

1^{st} Ionization Energy: energy needed to remove the first electron.

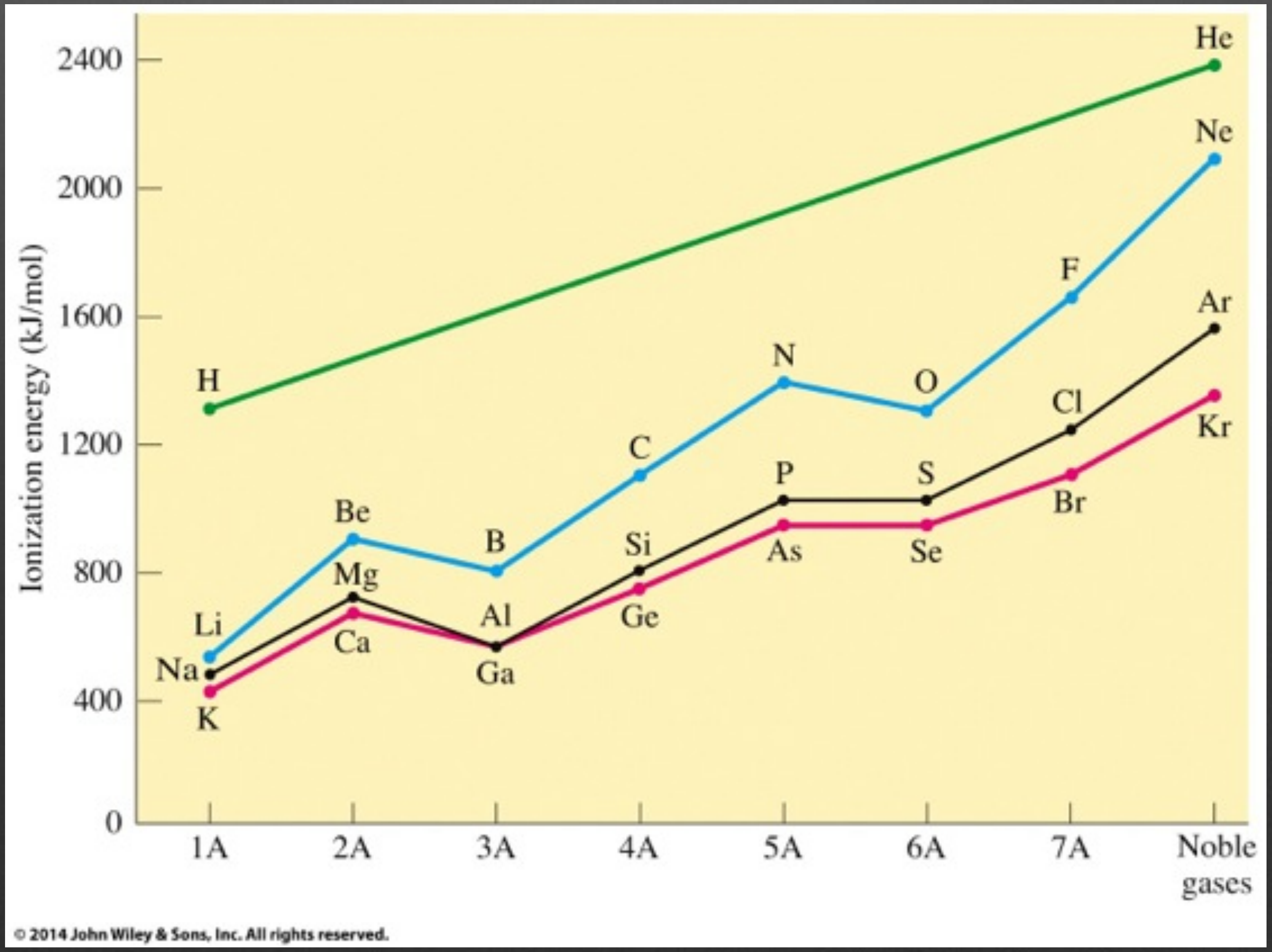
2^{nd} Ionization Energy: energy needed to remove the second electron.

Successive ionizations always increase in energy as the remaining electrons feel stronger attraction to the resulting cation produced.

Periodic Trends in Atomic Properties

Ionization Energy: generally decreases down a group and increases left to right across a period.

1st Ionization Energies for the Main Group Elements



Periodic Trends in Atomic Properties

Ionization Energy

Down a Group:
The electron removed is from a higher n level; farther from the nucleus, it feels less interaction, making it easier to remove.

Across a Period:
Left to right, n remains constant but atomic number increases; the increase in + nuclear charge creates a greater interaction with electrons, making it more difficult to remove outer electrons.

Periodic Trends in Atomic Properties

Ionization Energy

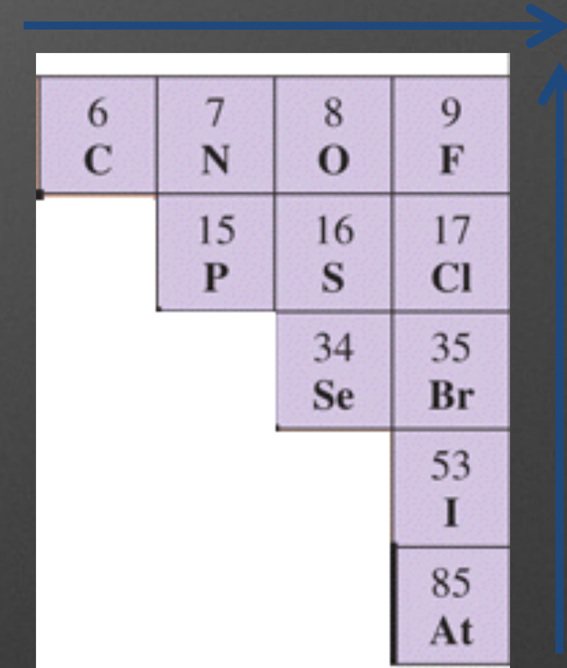
Extreme amounts of energy needed to remove electrons from noble gas configurations (in cyan).

Ionization Energies for Selected Elements

TABLE 11.1 Ionization Energies for Selected Elements*

Element	Required amounts of energy (kJ/mol) 1st e^-	2nd e^-	3rd e^-	4th e^-	5th e^-
H	1,314				
He	2,372	5,247			
Li	520	7,297	11,810		
Be	900	1,757	14,845	21,000	
B	800	2,430	3,659	25,020	32,810
C	1,088	2,352	4,619	6,222	37,800
Ne	2,080	3,962	6,276	9,376	12,190
Na	496	4,565	6,912	9,540	13,355

*Values are expressed in kilojoules per mole, showing energies required to remove 1 to 5 electrons per atom. Blue type indicates the energy needed to remove an electron from a noble gas electron structure.



Note: nonmetals generally have large ionization energies.

Nonmetals usually gain electrons to form anions. Most reactive nonmetals are in upper right corner of the periodic table.

Lewis Structures of Atoms

Valence (outermost) electrons in an atom are responsible for the formation of chemical bonds between atoms.

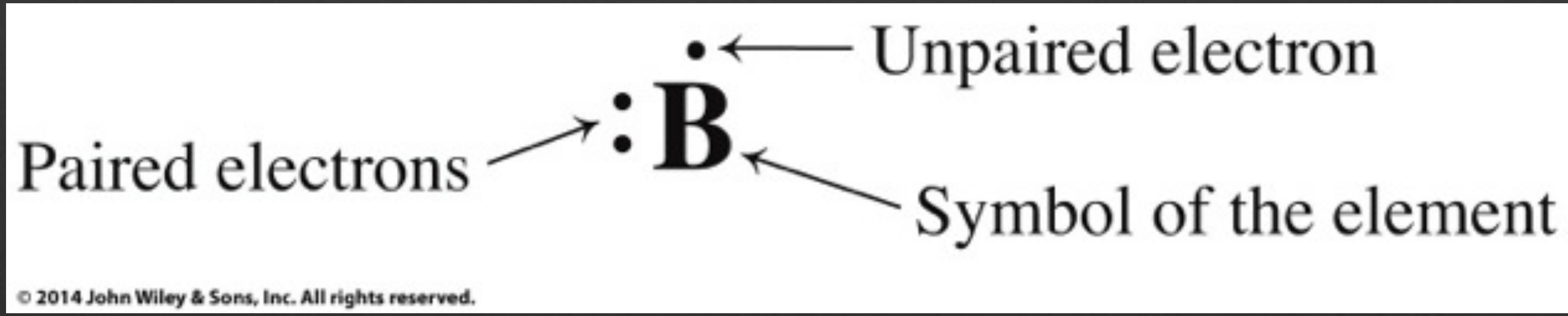
Lewis structures are a simple way of representing atoms and their valence electrons.

- uses dots to represent the valence electrons of an atom.

Example: Boron

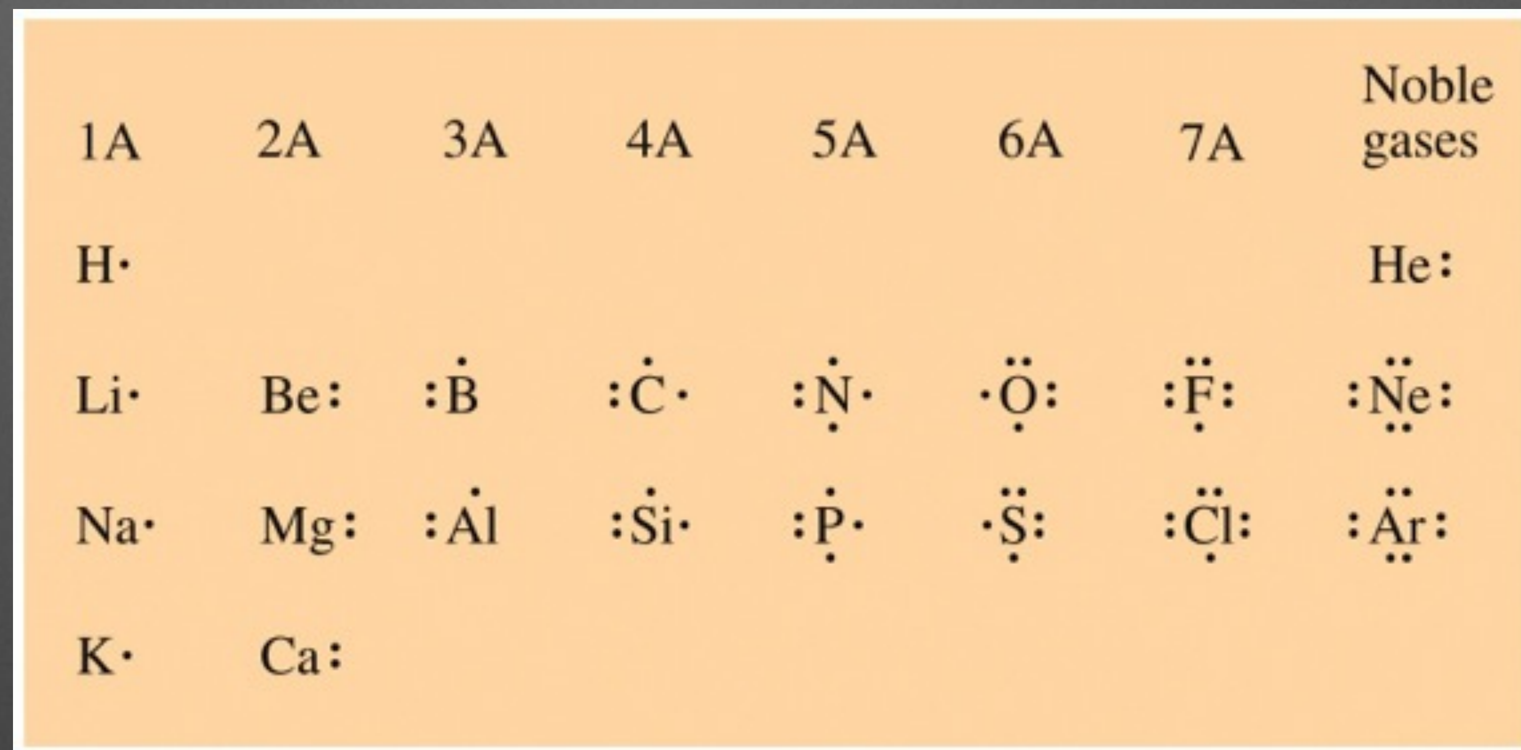
Electron Configuration: [He] $2s^2 2p^1$

Lewis structure of an atom



Lewis Structures of Atoms

Lewis Dot Structures of the First 20 Elements



Determining the Valence Electrons for an Atom:

For main group elements, the group number gives the number of valence electrons.

Example: Chlorine
(7 valence electrons because in Group 7A)

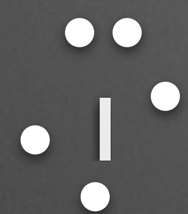
Lewis Structures of Atoms Practice

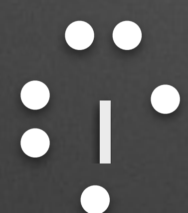
Draw the Lewis structure for the germanium atom.

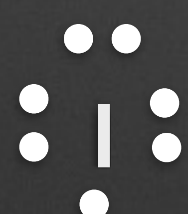
Ge is in Group 4A, so it contains 4 valence electrons.

Lewis Structures of Atoms Practice

Which of the following is the correct
Lewis structure for iodine?

a) 

b) 

c) 

I is in Group 7A, so it contains
7 valence electrons.

The Ionic Bond: Transfer of Electrons Between Atoms

Main group elements (Groups 1-7A) transfer electrons to attain an outer electron configuration that resembles stable noble gases.

For all elements of the second period and below, the most stable configuration consists of eight electrons in the valence energy level.

Valence electrons in the noble gases

TABLE 11.2 Arrangement of Electrons in the Noble Gases*

		Electron structure					
Noble gas	**Symbol**	$n = 1$	2	3	4	5	6
Helium	He	$1s^2$					
Neon	Ne	$1s^2$	$2s^22p^6$				
Argon	Ar	$1s^2$	$2s^22p^6$	$3s^23p^6$			
Krypton	Kr	$1s^2$	$2s^22p^6$	$3s^23p^63d^{10}$	$4s^24p^6$		
Xenon	Xe	$1s^2$	$2s^22p^6$	$3s^23p^63d^{10}$	$4s^24p^64d^{10}$	$5s^25p^6$	
Radon	Rn	$1s^2$	$2s^22p^6$	$3s^23p^63d^{10}$	$4s^24p^64d^{10}4f^{14}$	$5s^25p^65d^{10}$	$6s^26p^6$

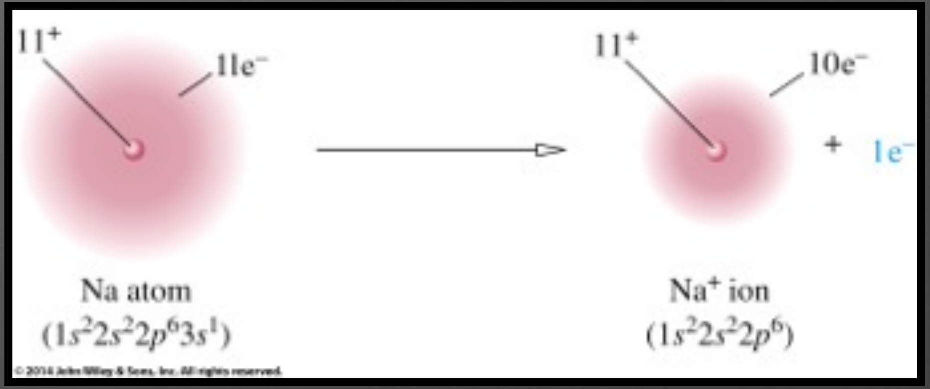
*Each gas except helium has eight electrons in its outermost energy level.

Ionic Compound Formation

Consider Na and Cl.

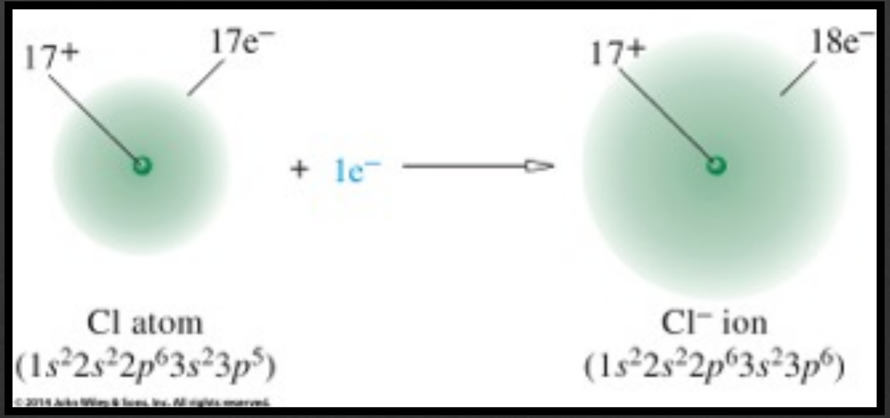
Na has 11 total electrons. ($1s^2 2s^2 2p^6 3s^1$)

To achieve a noble gas core, it is easiest to lose one electron and become a cation.



Cl has 17 total electrons. ($1s^2 2s^2 2p^6 3s^2 3p^5$)

To achieve a noble gas core, it is easiest to gain one electron and become an anion.



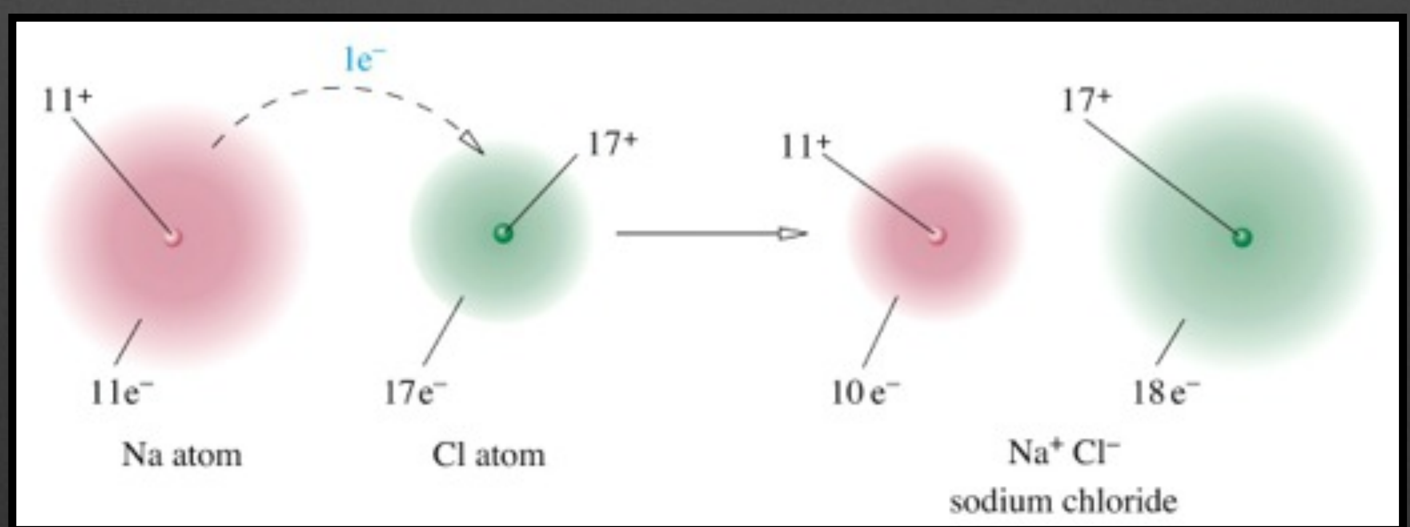
Ionic Compound Formation

Bonding between Na and Cl.

If sodium transfers an electron to Cl, both become ions and achieve a noble gas configuration.

The electrostatic attraction holding the two ions together is an ionic bond.

Ionic bond: attraction between oppositely charged ions.

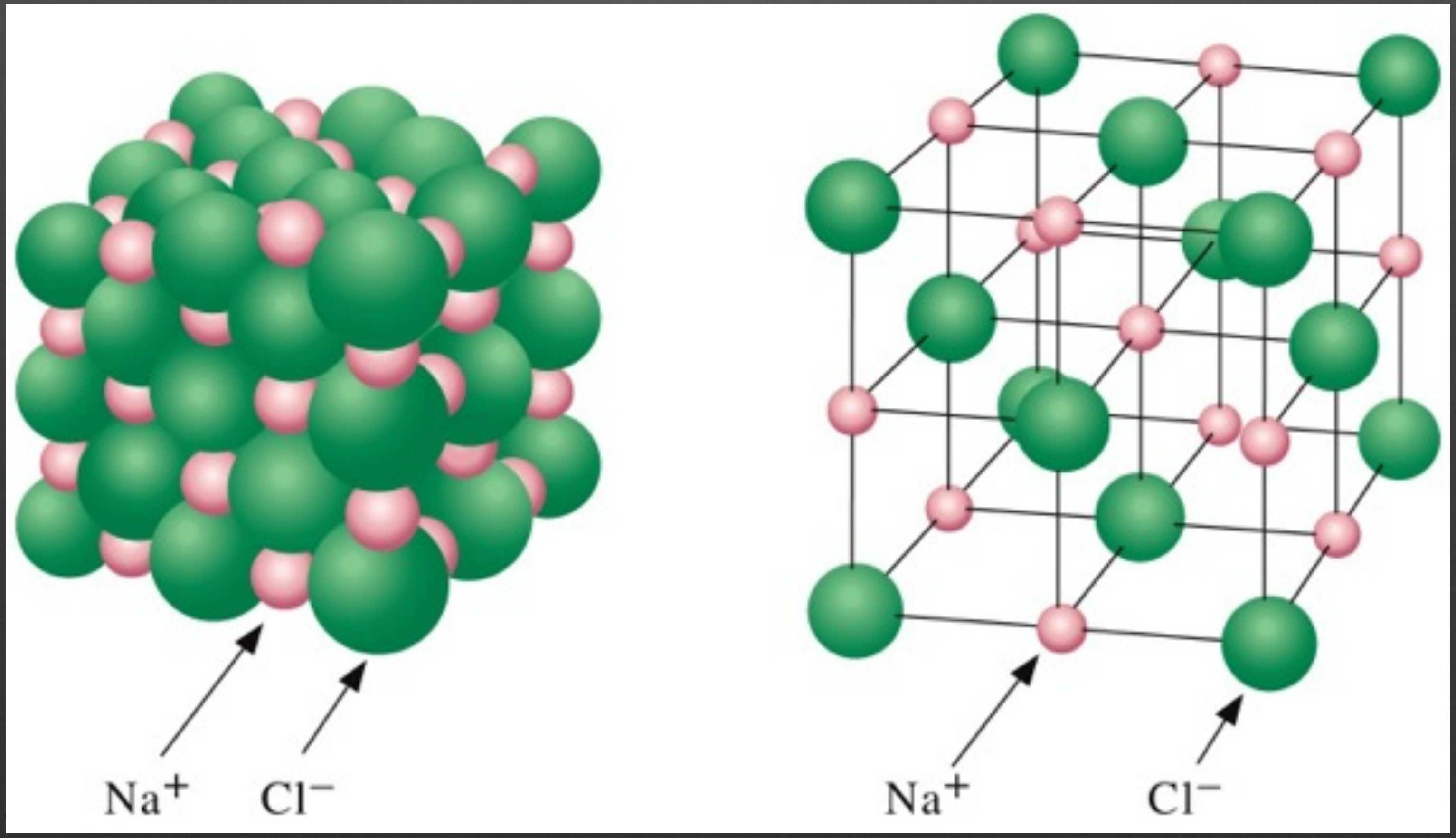


$$\mathrm{Na\cdot} + \mathrm{:\ddot{\underset{..}{Cl}}:} \longrightarrow [\mathrm{Na}]^{+}\left[\mathrm{:\ddot{\underset{..}{Cl}}:}\right]^{-}$$

Ionic Compound Structure

In the solid state, NaCl packs in an ordered three dimensional cubic array. Each Na+ is surrounded by 6 Cl– anions. In turn, each Cl– is surrounded by 6 Na+ cations.

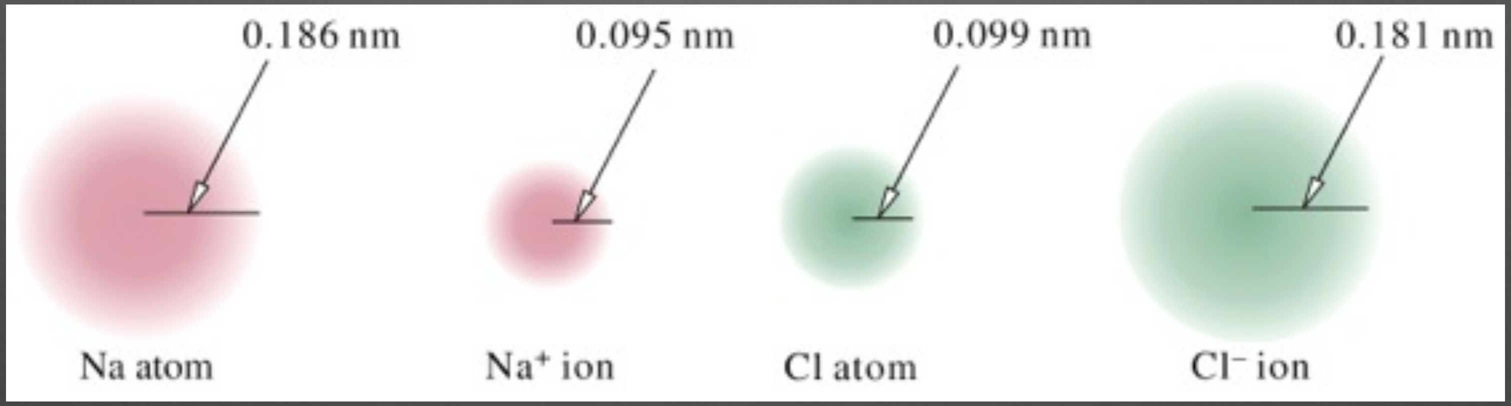
Lattice Diagram for NaCl



Cation/Anion Size

Cations are smaller than the parent atom while anions are larger than the parent atom.

Relative radii of Na and Cl atoms and their corresponding ions.



In NaCl, Na+ is much smaller than Na, as there is greater attraction between the nucleus and the 10 electrons remaining relative to the parent atom.

Cl– is larger than Cl because the attraction between the nucleus is reduced by adding an extra electron to make an anion.

Cation/Anion Size

Dramatic size changes from the free atoms are typically observed.

Atomic Radii of Selected Metals and Nonmetals

TABLE 11.3 Change in Atomic Radii (nm) of Selected Metals and Nonmetals*

Atomic radius		Ionic radius		Atomic radius		Ionic radius	
Li	0.152	Li^{+}	0.060	F	0.071	F^{-}	0.136
Na	0.186	Na^{+}	0.095	Cl	0.099	Cl^{-}	0.181
K	0.227	K^{+}	0.133	Br	0.114	Br^{-}	0.195
Mg	0.160	Mg^{2+}	0.065	O	0.074	O^{2-}	0.140
Al	0.143	Al^{3+}	0.050	S	0.103	S^{2-}	0.184

*The metals shown lose electrons to become positive ions. The nonmetals gain electrons to become negative ions.

Ionic Bonding

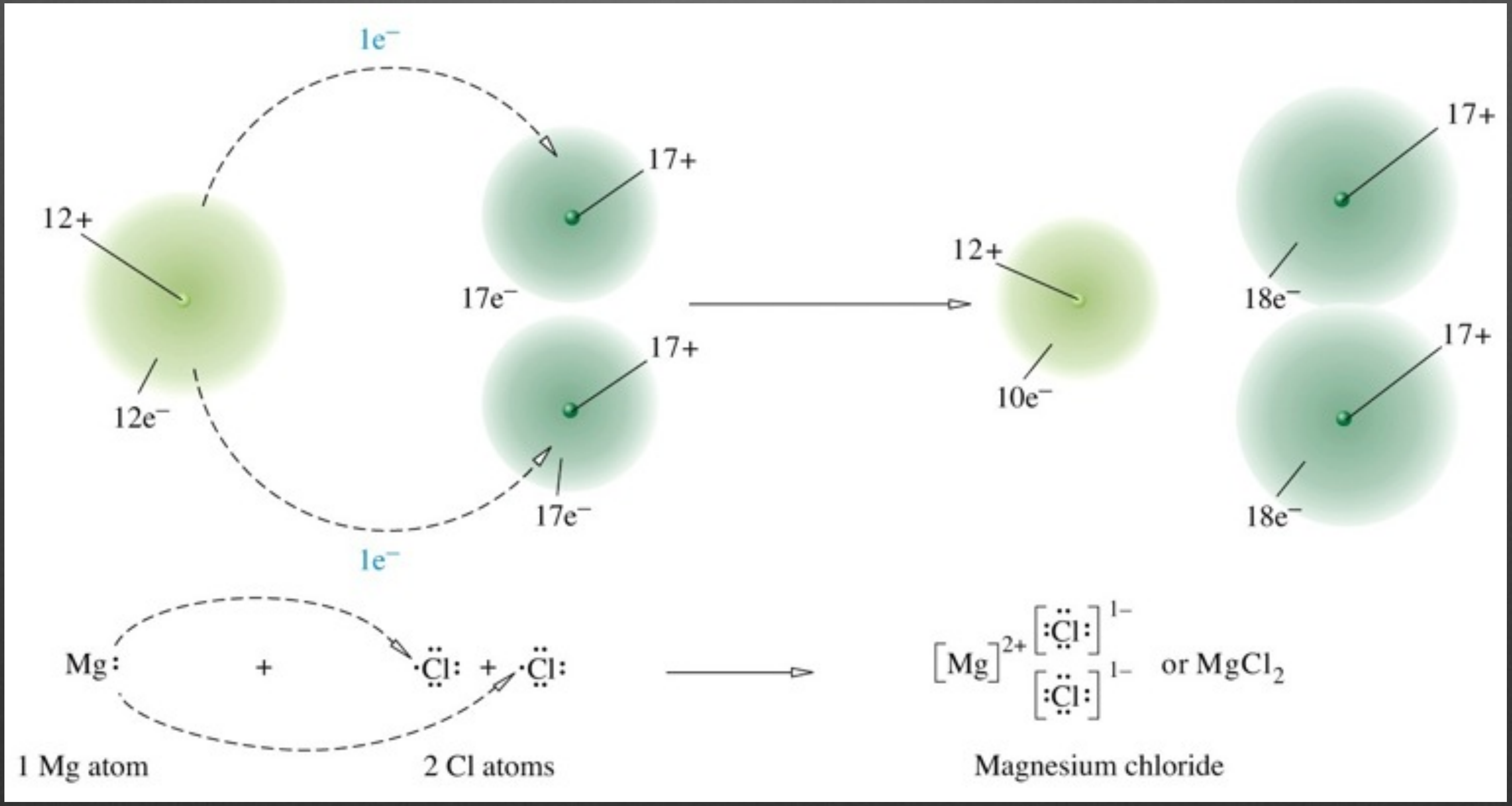
More than one electron can be transferred between a metal and a nonmetal to form ionic compounds.

Mg could lose 2 electrons to achieve a noble gas configuration ([Ne]$3s^2$).

But one Cl needs to gain only 1 electron to achieve a noble gas core ([Ne]$3s^2 3p^5$).

Ionic Bonding

The two electrons from Mg are then transferred to two different Cl atoms.



Ionic Bonding

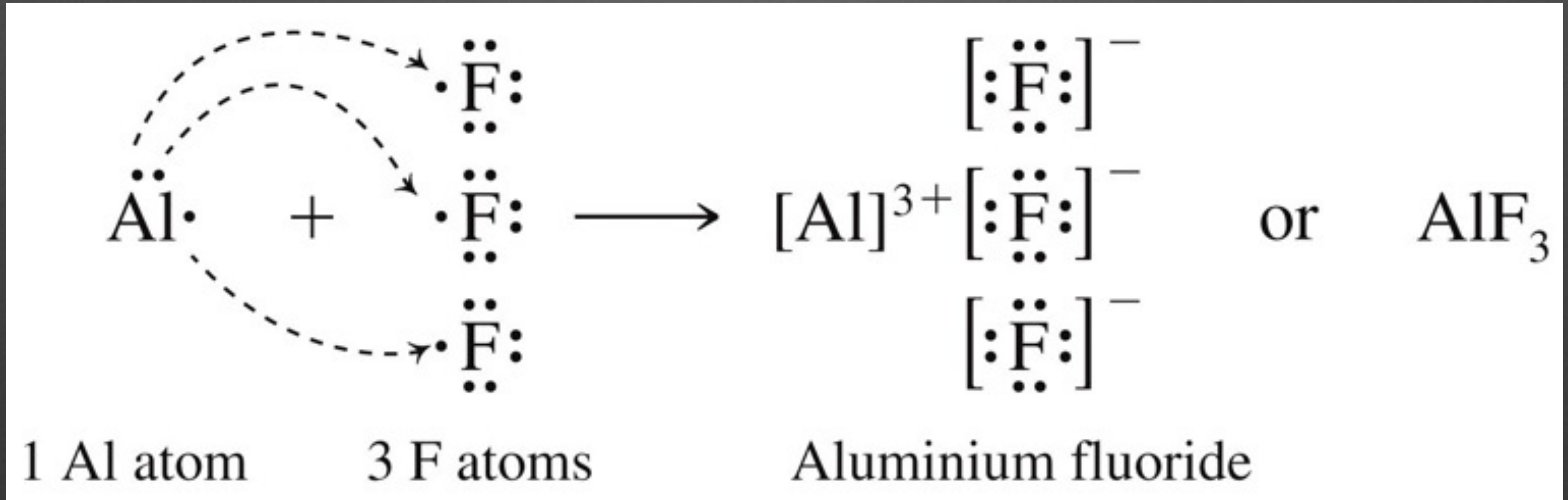
More than one electron can be transferred between a metal and a nonmetal to form ionic compounds.

Al could lose 3 electrons to achieve a noble gas configuration ([Ne]$3s^2 3p^1$).

But one F needs to gain only 1 electron to achieve a noble gas core ([He]$2s^2 2p^5$).

Ionic Bonding

Three electrons need to be transferred; one to each fluorine.



Predicting Formulas of Ionic Compounds

In almost all stable compounds of main group elements, all atoms attempt to attain noble gas configuration.

Metals lose electrons to do so; nonmetals gain electrons.

Ionic compounds are charge neutral.
(The overall + and – charges must balance!)

Example: What is the formula containing Ba and S?

Ba is in Group 2, would like to form Ba^{2+}
S is in Group 6, would like to form S^{2-}

Charge balance of Ba^{2+} and S^{2-} = **BaS**

Predicting Formulas of Ionic

Because atoms in the same groups have similar valence configurations, they typically form the same types of ions and compounds with the same general formulas.

TABLE 11.4 Formulas of Compounds Formed by Alkali Metals

Lewis structure	Oxides	Chlorides	Bromides	Sulfates
Li·	Li_2O	LiCl	LiBr	Li_2SO_4
Na·	Na_2O	NaCl	NaBr	Na_2SO_4
K·	K_2O	KCl	KBr	K_2SO_4
Rb·	Rb_2O	RbCl	RbBr	Rb_2SO_4
Cs·	Cs_2O	CsCl	CsBr	Cs_2SO_4

This provides predictive power to the periodic table!

Predicting Formulas of Ionic Compounds

What is the formula for potassium phosphide?

K is in Group 1, would like to form K^+
P is in Group 5, would like to form P^{3-}

Charge balance of K^+ and P^{3-} = K_3P

What is the formula for sodium sulfide?

Na is in Group 1, would like to form Na^+
S is in Group 6, would like to form S^{2-}

Charge balance of Na^+ and S^{2-} = Na_2S

Predicting Formulas of Ionic Compounds

What is the formula for magnesium nitride?

a. MgN

b. Mg_2N

c. Mg_3N_2

d. Mg_2N_3

Mg is in Group 2, would like to form Mg^{2+}

N is in Group 5, would like to form N^{3-}

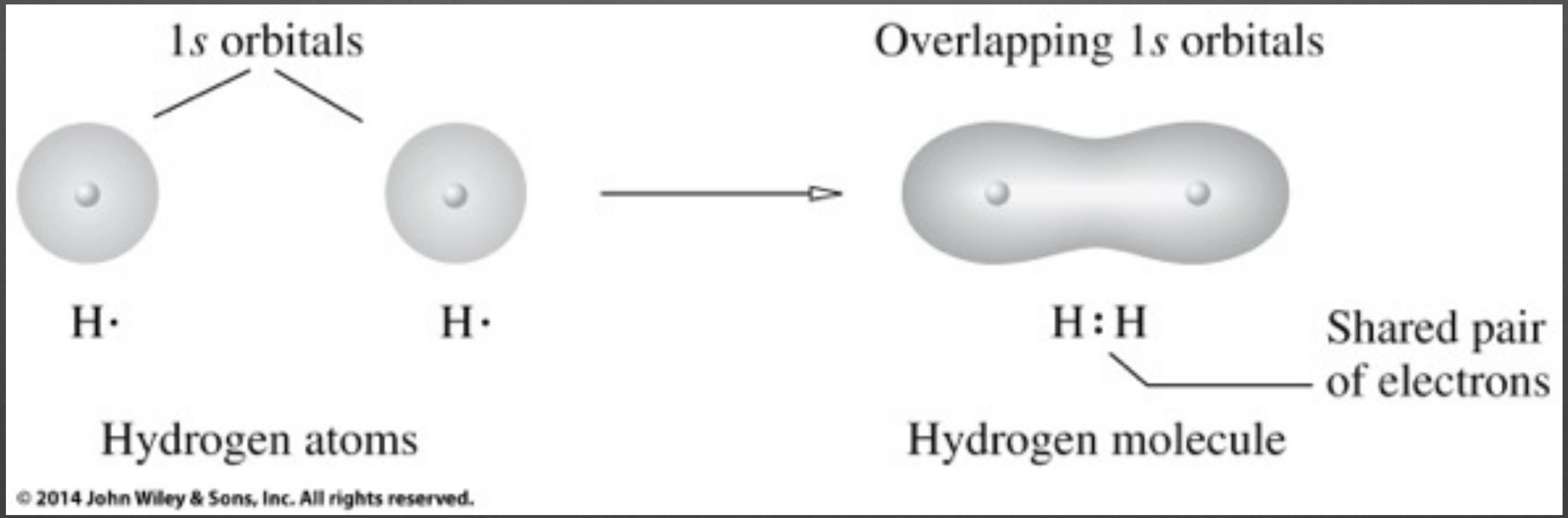
Charge balance of Mg^{2+} and N^{3-} = Mg_3N_2

Covalent Bonds

Molecules are discrete units held together by covalent bonds.

Covalent bond: pair of electrons shared between 2 atoms

Example: the covalent bond in H_2



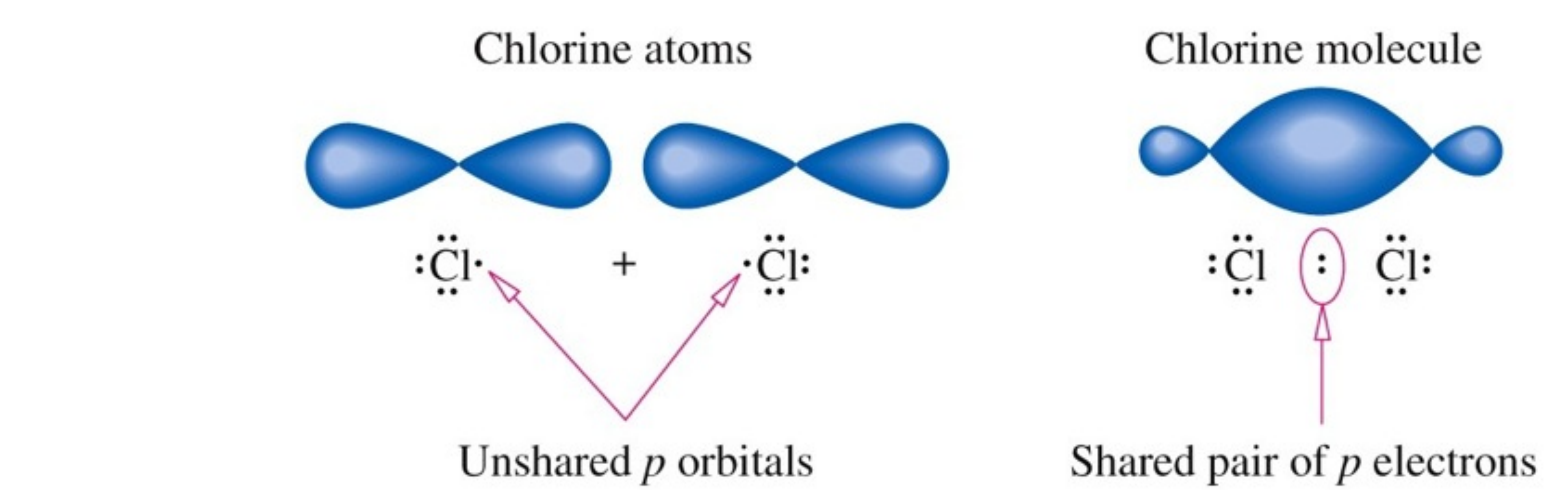
Formation of the H_2 molecule from two hydrogen atoms. The 1s orbitals on each atom overlap, forming a covalent bond between the atoms.

Each atom contributes one electron to the bond.

Covalent Bonds

The bonding in Cl2

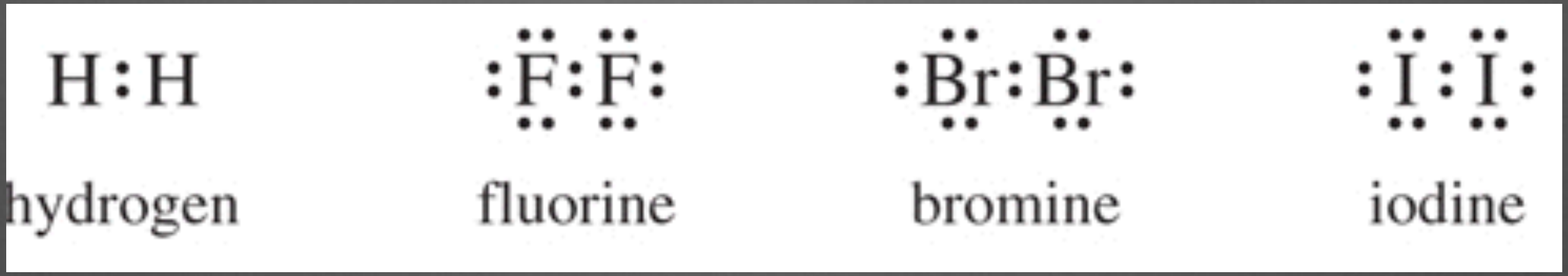
Overlap of a 3p orbital on each chlorine gives rise to a bond between chlorine atoms.



Each atom contributes one electron to the bond.

Covalent Bonds Between Other Elements

Single bonds are typically formed between hydrogen or halogens because each element needs 1 electron to achieve a noble gas configuration.



Typically, shared electrons between atoms are replaced with a dashed line (—).

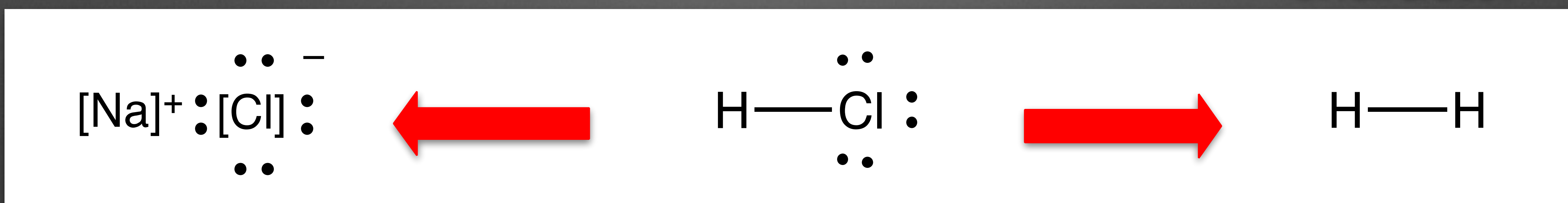
Atoms, to be more stable, can also form 2 or 3 covalent bonds as needed.

Bonding: A Continuum

Ionic and covalent bonding represent two extremes. Electrons are either fully transferred between atoms (ionic) or shared (covalent).

Polar covalent

Purely covalent character



Purely ionic character

Electron transfer

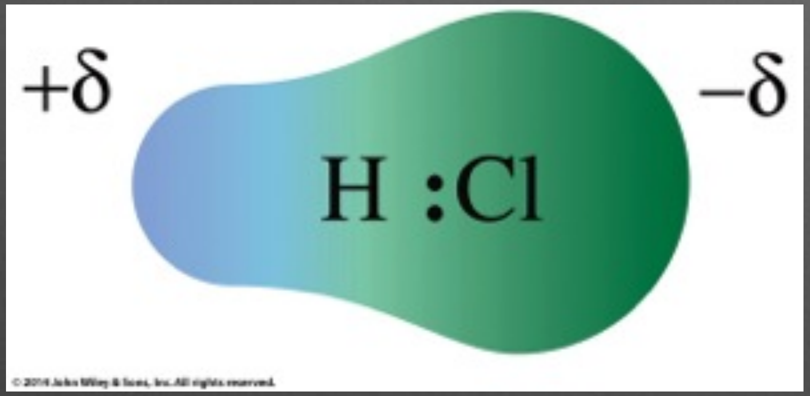
Electron sharing

Realize bonding in between the extremes also exists (i.e. polar covalent bonds).

Electronegativity

Electronegativity: a measure of the ability of an atom to attract electrons in a covalent bond.

Example



Chlorine is more electronegative than hydrogen.

The electrons shared between H and Cl remain closer to Cl than H as a result. This gives the Cl a partial (δ) negative charge. H then has a partial positive charge.

Electronegativity Scale

The most electronegative element, F, is assigned a value of 4.0.

The higher the electronegativity, the stronger an atom attracts electrons in a covalent bond.

Bond polarity is determined by the difference in electronegativity of the two atoms sharing electrons.

Electronegativity Scale

General Trends in Electronegativity

Increases left to right across the periodic table.
Decreases down a main group.

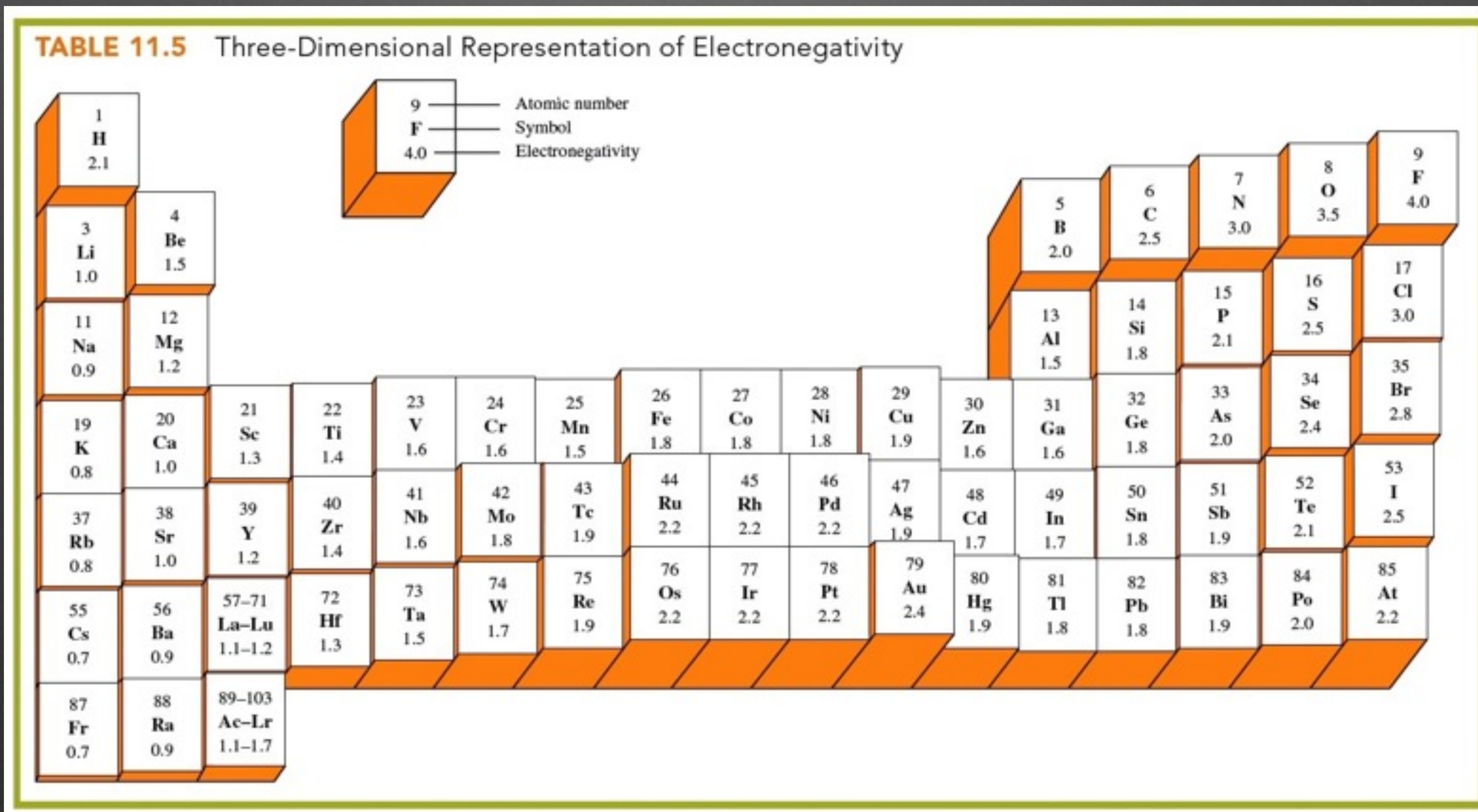


TABLE 11.5 Three-Dimensional Representation of Electronegativity

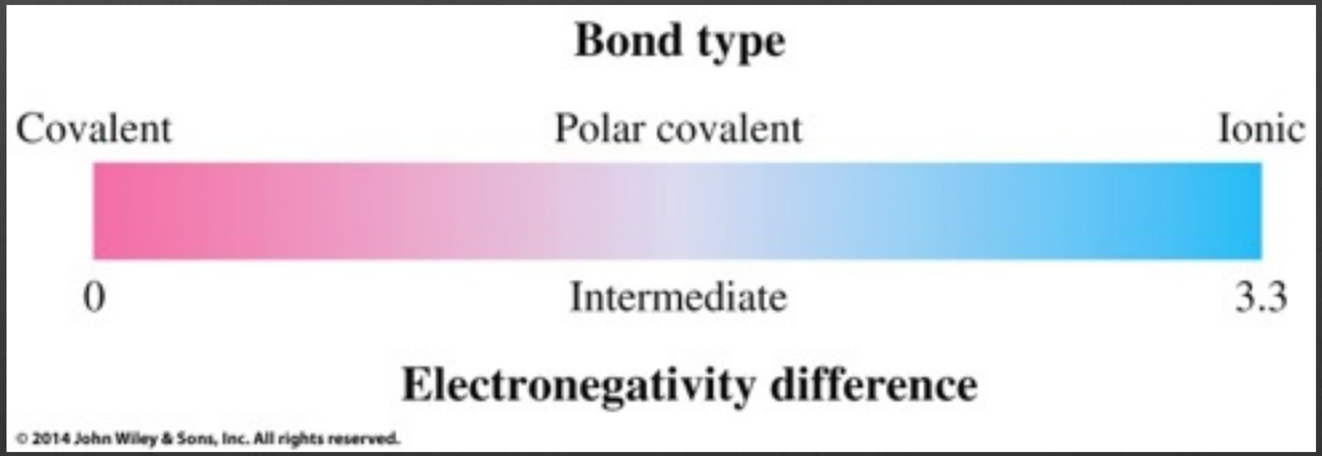
Bonding Continuum

Bond polarity is determined by the difference in electronegativity of the two atoms sharing electrons.

If the atoms in the covalent bond are the same, the bond is nonpolar and the electrons are shared equally.

If the difference in electronegativities is > 2, the bonding is considered ionic.

If the difference in electronegativties is < 2, the bonding is considered polar covalent (unequal sharing).



Bonding Continuum: Nonpolar Bonds

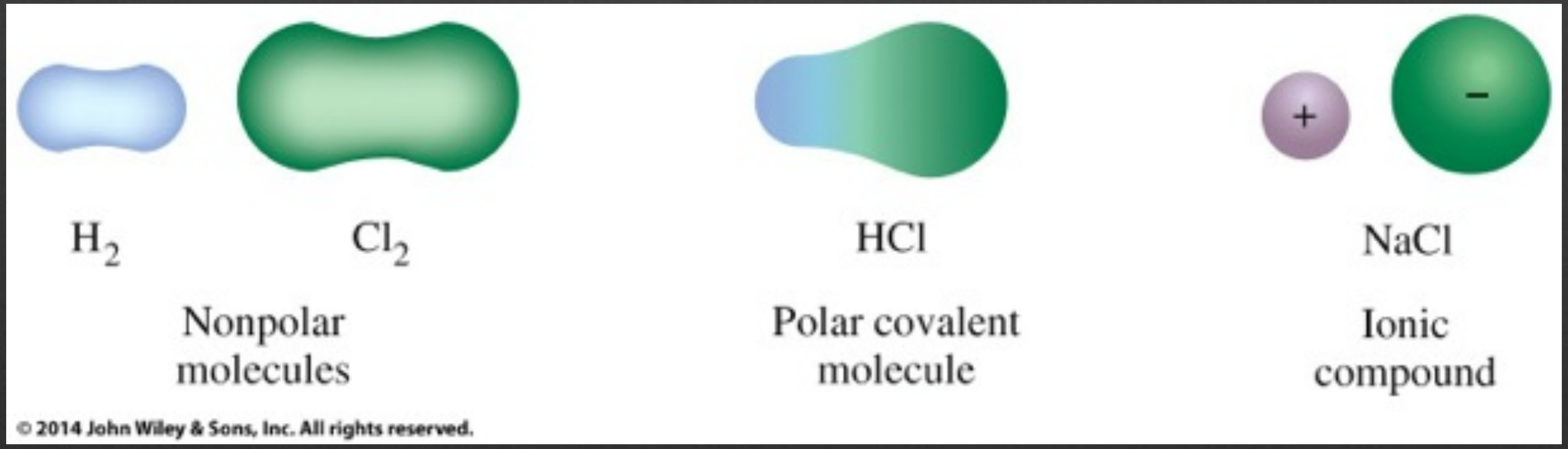
Nonpolar covalent bonds have very small or no difference in electronegativities between the two atoms in the bond.

Electrons are shared equally between atoms.

Examples:

C-S bonds: electronegativity difference = 2.5 - 2.5 = 0

N-Cl bonds: electronegativity difference = 3.0 - 3.0 = 0



Bonding Continuum: Polar Covalent Bonds

Polar covalent bonds occur when two atoms share electrons unequally.

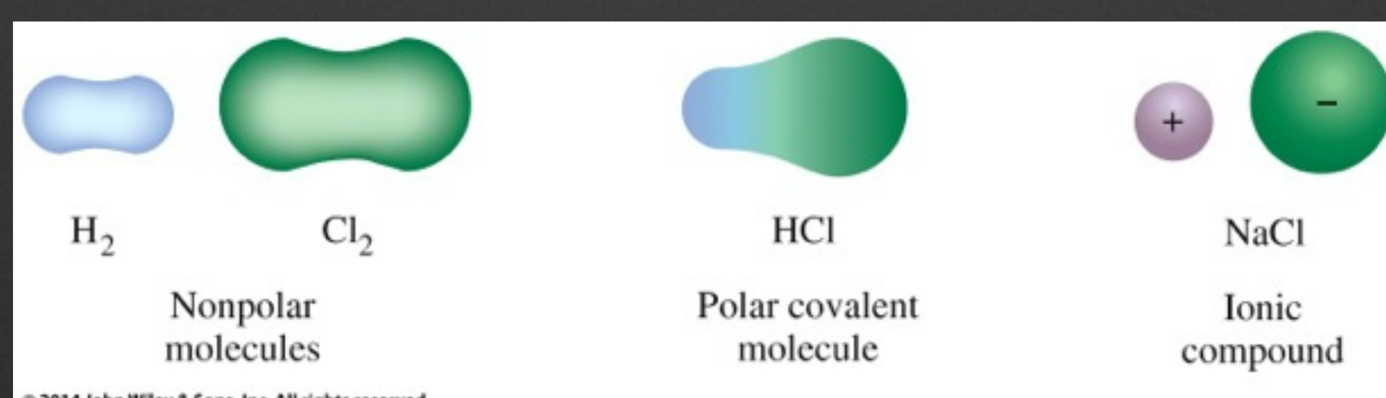
Examples:

P-O bonds: electronegativity difference = 3.5 -2.1 = 1.4

P  O

N-C bonds: electronegativity difference = 3.0 -2.5 = 0.5

N  C



Bonding Continuum Practice

A polar covalent bond would form between which two atoms?

a. Be and F

b. H and Cl

c. Na and O

d. F and F

Examine the electronegativity difference between each set of atoms

Be/F = 4.0 – 1.5 = 2.5
H/Cl = 3.0 – 2.1 = 0.9
Na/O = 3.5 – 0.9 = 2.6
F/F = 4.0 – 4.0 = 0.0

Only the HCl bond is within the range of polar covalent interactions (0-2).

Polar Bonds and Polar Molecules

A molecule can contain polar bonds (if the two atoms are different) but still be nonpolar, depending on the geometry of the molecule.

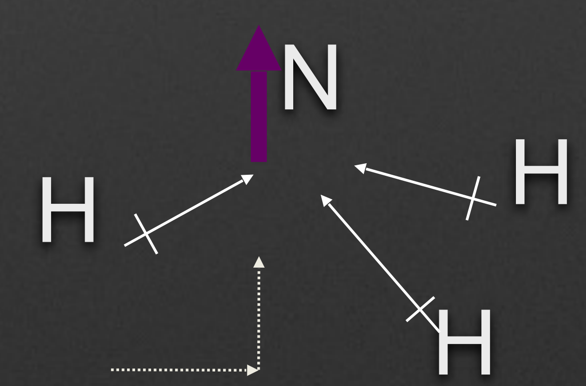
Example: Carbon dioxide

Though the molecule contains polar C=O bonds, the molecule is linear, so the C-O dipoles cancel each other.

$$O = C = O$$

Symmetric arrangements of polar bonds result in nonpolar molecules.

Asymmetric arrangements of polar bonds result in polar molecules.



Lewis Structures of Compounds

Lewis structures are a convenient way to display the bonding in molecules or ions.

General Strategies for Writing Lewis Structures

1. Determine the total number of valence electrons available for bonding by summing the valence electrons for all the atoms. Charges also affect the number of electrons available!

2. Draw the skeletal arrangement of the atoms by connecting each with a single bond. (H can only form one bond and can never be a central atom!)

Lewis Structures of Compounds

General Strategies for Writing Lewis Structures

3. Subtract the number of bonding electrons from the total available.

4. Add electron pairs to each atom to complete their octets as needed.

5. Form double or triple bonds as needed to complete octets around atoms as needed.

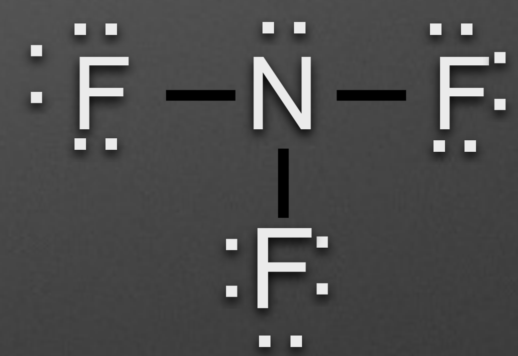
Lewis Structures of Compounds

Draw the Lewis structure for NF_3.

Sum of valence electrons: $5 + 3(7) = 26$

Remember the valence electrons around an atom are equal to their group number for main group elements.

Arrange the skeletal structure of the molecule.



Subtract the bonding electrons from the total: $26 – 6 = 20$

Add lone pairs to complete each atoms’ octet.

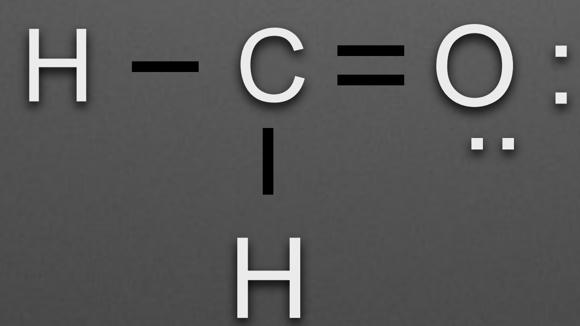
A molecule’s shape is not predicted by a Lewis structure.

Lewis Structures of Compounds

Draw the Lewis structure for CH_2O.

Sum of valence electrons: 4 + 2(1) + 6 = 12

Arrange the skeletal structure of the molecule.



Subtract the bonding electrons from the total: 12 – 6 = 6

Add lone pairs to complete each atoms’ octet.

Form double or triple bonds as needed to complete octets.

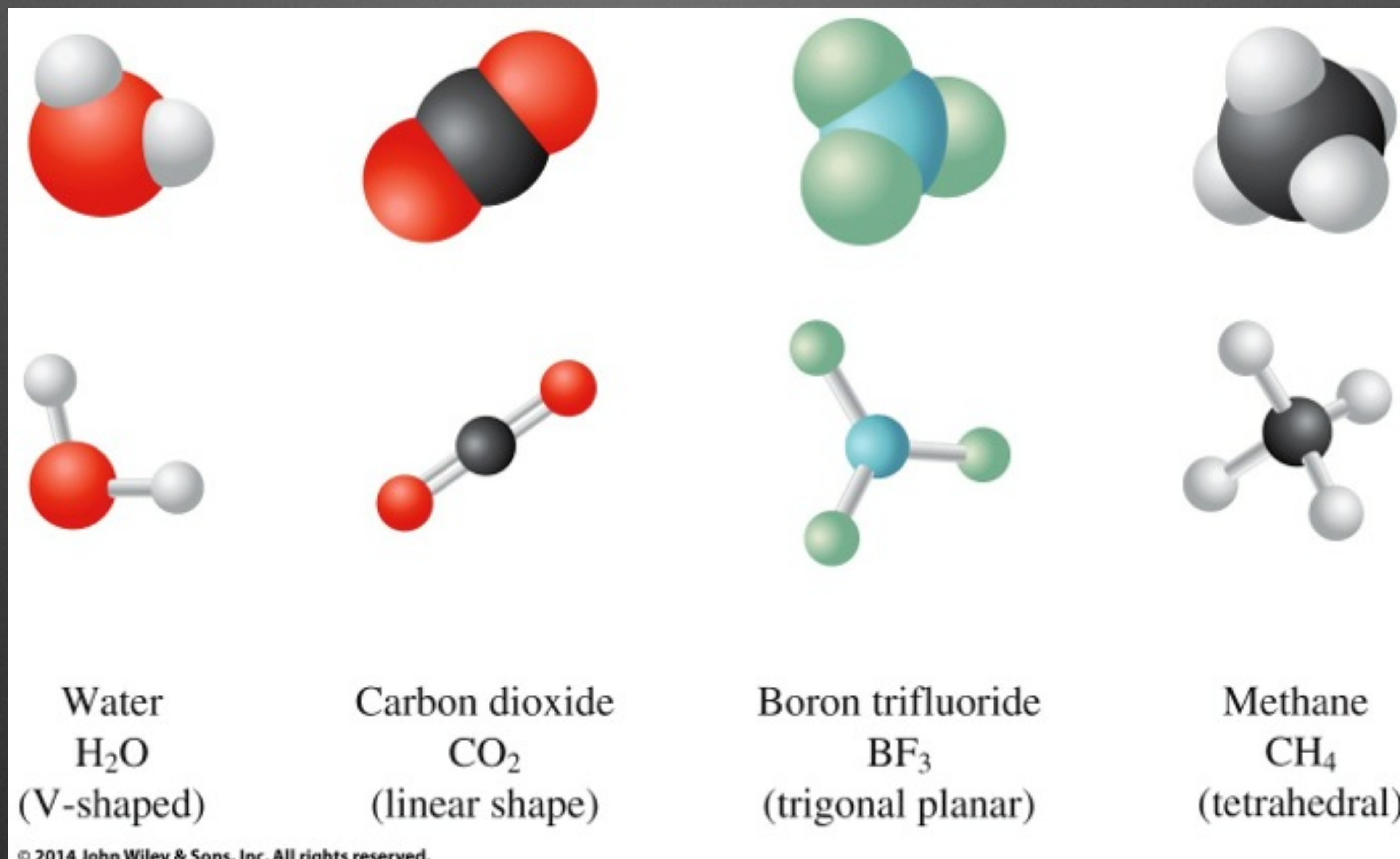
VSEPR Theory

Valence Shell Electron Pair Repulsion Theory: Theory developed to help predict the geometry of molecules based on the bonding and nonbonding electrons around a central atom.

Molecules adopt geometries to minimize the electrostatic repulsion between electron pairs.

Molecular Shape

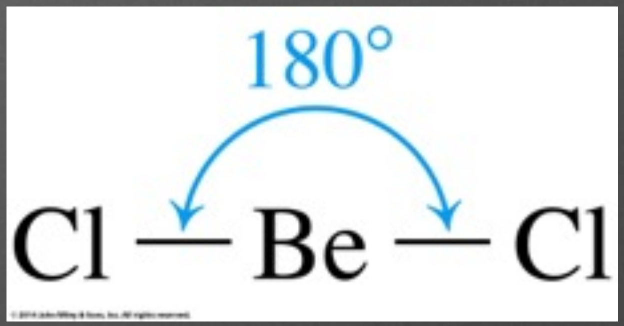
Shape of some common molecules.
Each molecule is represented by both a ball and stick model (bottom) and a space filling model (top).



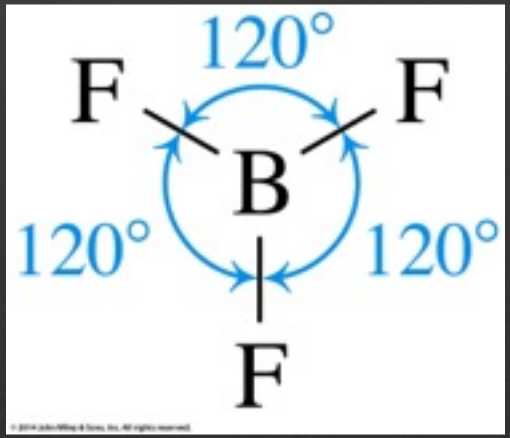
VSEPR Theory

A linear geometry results when two electron pairs surround a central atom.

A linear geometry maximizes distance of the bonding pairs.

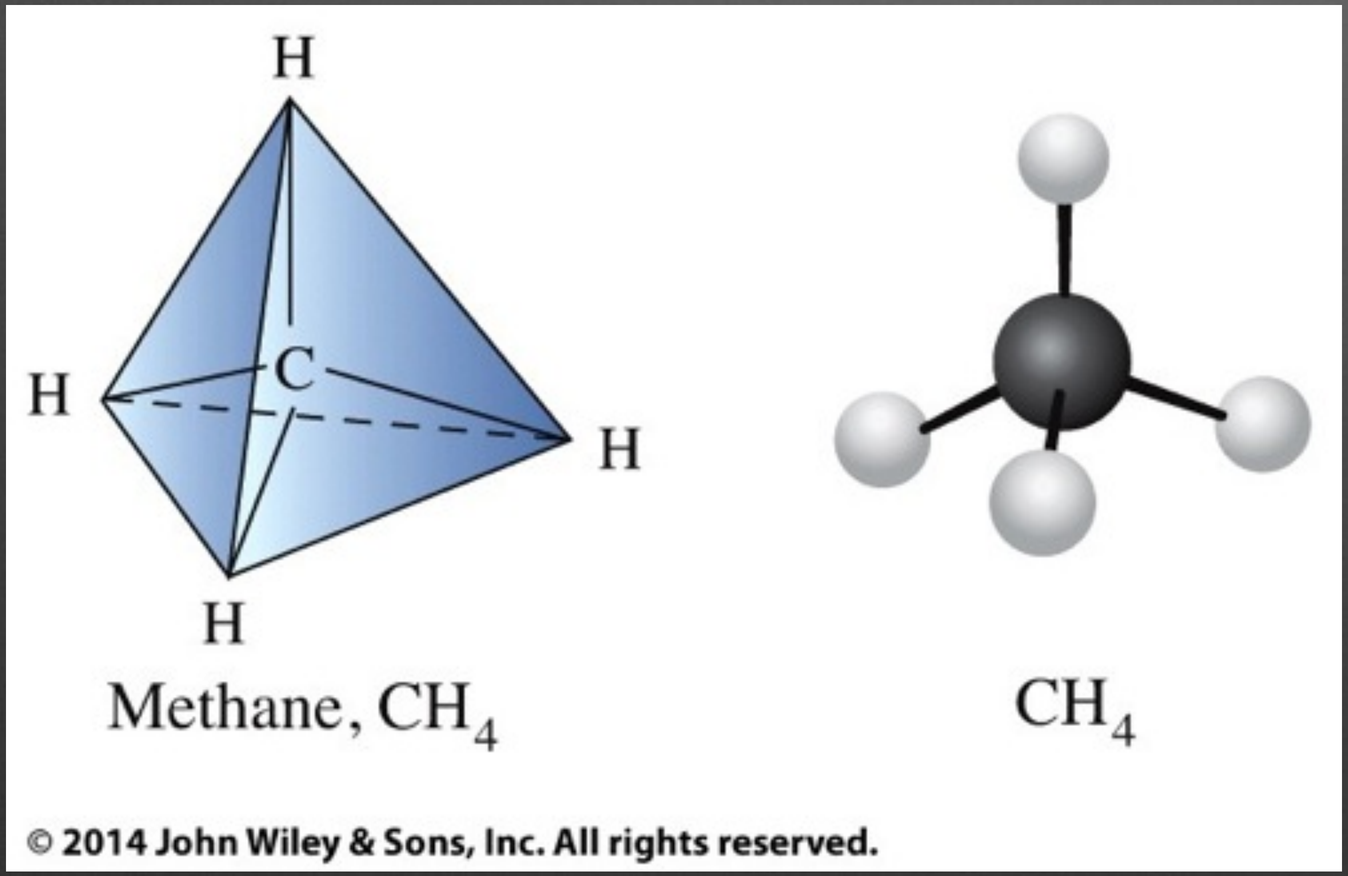


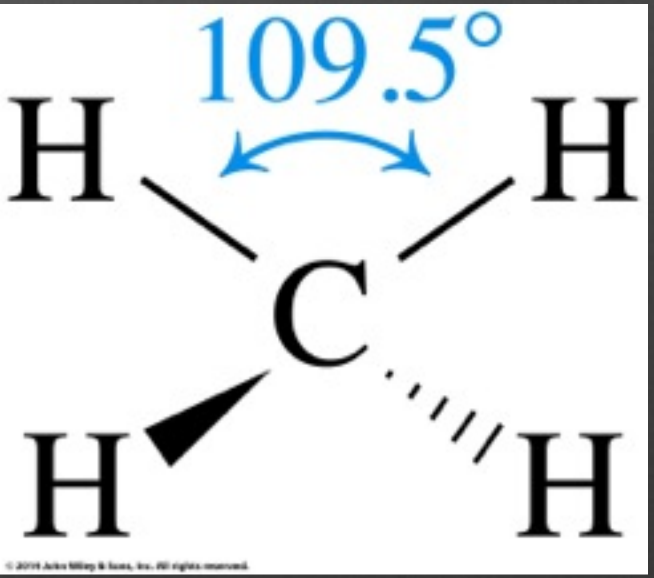
A trigonal planar geometry results when three electron pairs surround a central atom.



VSEPR Theory

A tetrahedral geometry results when four electron pairs surround a central atom.



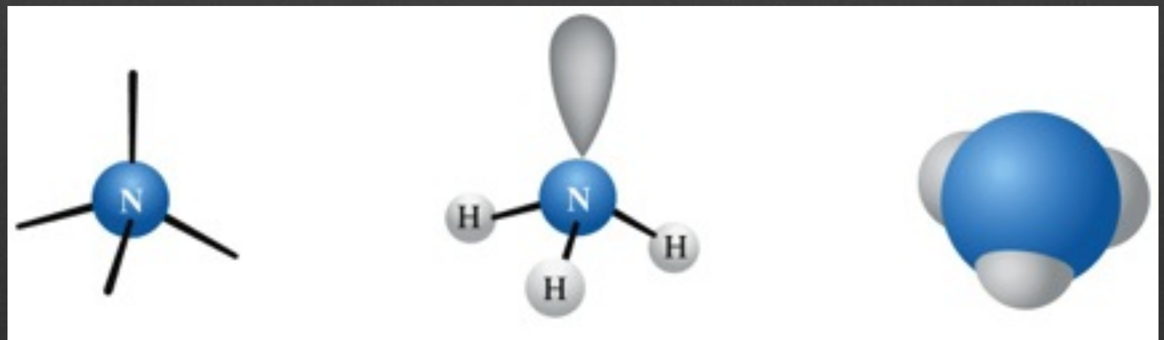


The Effect of Lone Pairs on Shape

Consider ammonia (NH_3) which contains 4 electron pairs (3 in bonds, 1 as a lone pair on N).

The electron pair geometry (based on 4 pairs) is tetrahedral (a & b).

The geometry considering only the bonding electrons is pyramidal (c).



(a) (b) (c)

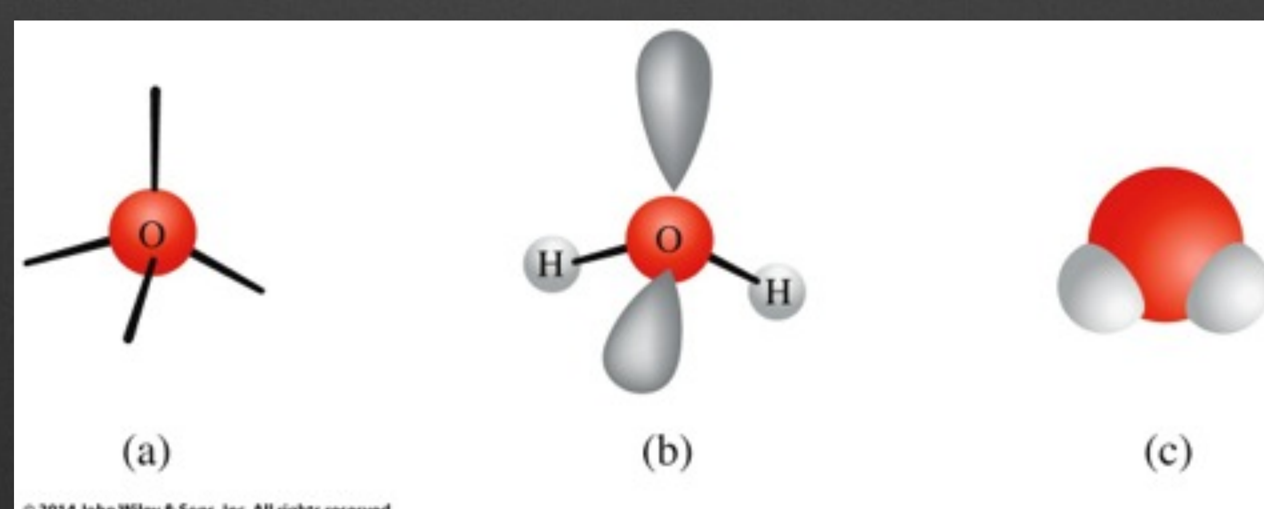
The Effect of Lone Pairs on Shape

Consider water (H_2O) which contains 4 electron pairs (2 in bonds, 2 as lone pairs on O).

The electron pair geometry (based on 4 pairs) is tetrahedral (a & b).

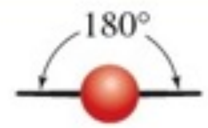
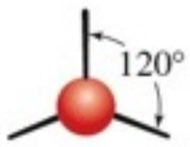
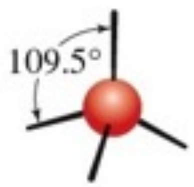
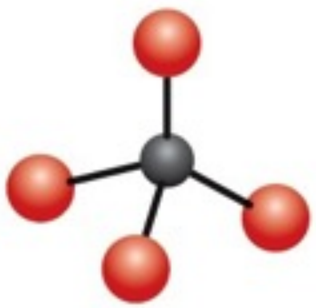
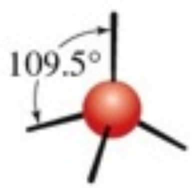
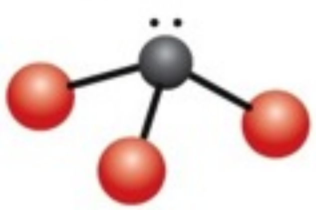
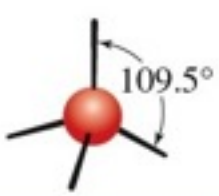
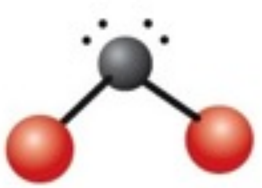
$$H - \underset{\cdot\cdot}{\ddot{O}} - H$$

The geometry considering only the bonding electrons is bent (c).



The Effect of Lone Pairs on Shape

TABLE 11.6 Arrangement of Electron Pairs and Molecular Structure

Number of electron pairs	Electron pair arrangement	Ball-and-stick model	Bonds	Molecular structure	Molecular structure model
2	Linear		2	Linear	
3	Trigonal planar		3	Trigonal planar	
4	Tetrahedral		4	Tetrahedral	
4	Tetrahedral		3	Pyramidal	
4	Tetrahedral		2	Bent	

Determining Molecular Shape

1. Draw the Lewis structure for a molecule.

2. Count the electron pairs and arrange them to minimize electrostatic repulsion.

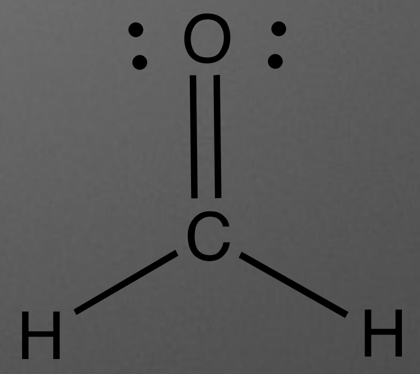
3. Determine the position of atoms based on 2.

4. Provide the molecular geometry for the molecule, which results from considering only bonding electrons.

Determining Molecular Shape Practice

What is the molecular geometry for formaldehyde, CH_2O?

1. Draw the Lewis structure for a molecule.



a. linear

b. trigonal planar

c. bent

d. tetrahedral

2. Count the electron pairs and arrange them to minimize electrostatic repulsion.
3. Determine the position of atoms based on 2.
4. Provide the molecular geometry for the molecule, which results from considering only bonding electrons.

Determining Molecular Shape Practice

What is the molecular geometry for PH_3?

a. linear

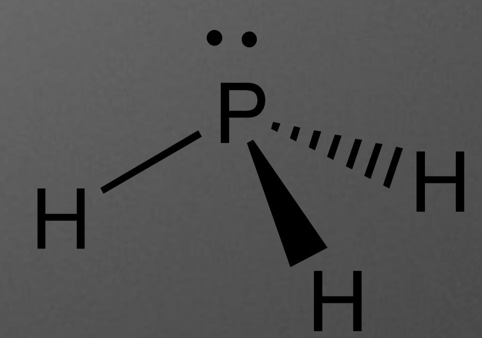
b. trigonal planar

c. bent

d. pyramidal

e. tetrahedral

1. Draw the Lewis structure for a molecule.



2. Count the electron pairs and arrange them to minimize electrostatic repulsion.
3. Determine the position of atoms based on 2.
4. Provide the molecular geometry for the molecule, which results from considering only bonding electrons.

Determining Molecular Shape Practice

What is the molecular geometry for CCl_4?

a. linear

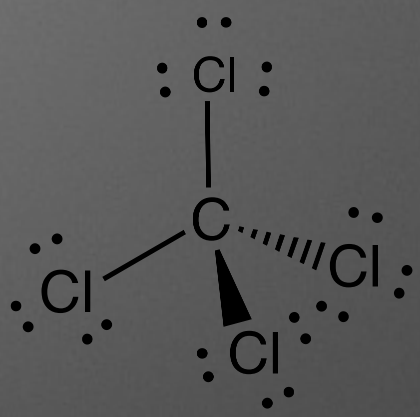
b. trigonal planar

c. bent

d. trigonal pyramidal

e. tetrahedral

1. Draw the Lewis structure for a molecule.



2. Count the electron pairs and arrange them to minimize electrostatic repulsion.
3. Determine the position of atoms based on 2.
4. Provide the molecular geometry for the molecule, which results from considering only bonding electrons.

Determining Molecular Shape Practice

Is CCl_4 polar or nonpolar?

a. polar due to polar bond arranged symmetrically about C.

b. polar due to polar bonds arranged asymmetrically about C.

c. nonpolar due to polar bonds being arranged symmetrically about C.

d. nonpolar as all bonds in the molecule are not polarized.