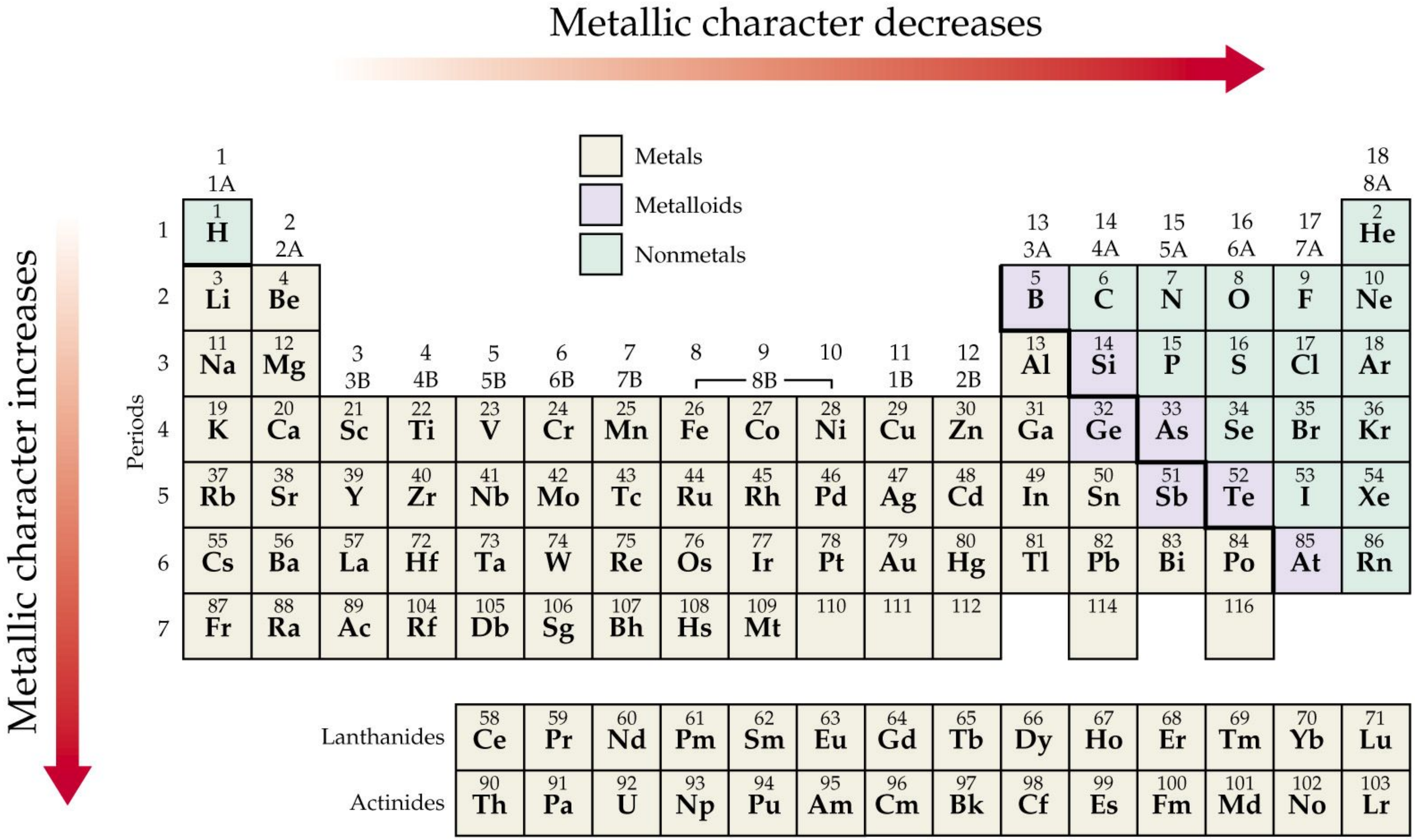
Chemical Bonds: The Formation of Compounds from Atoms

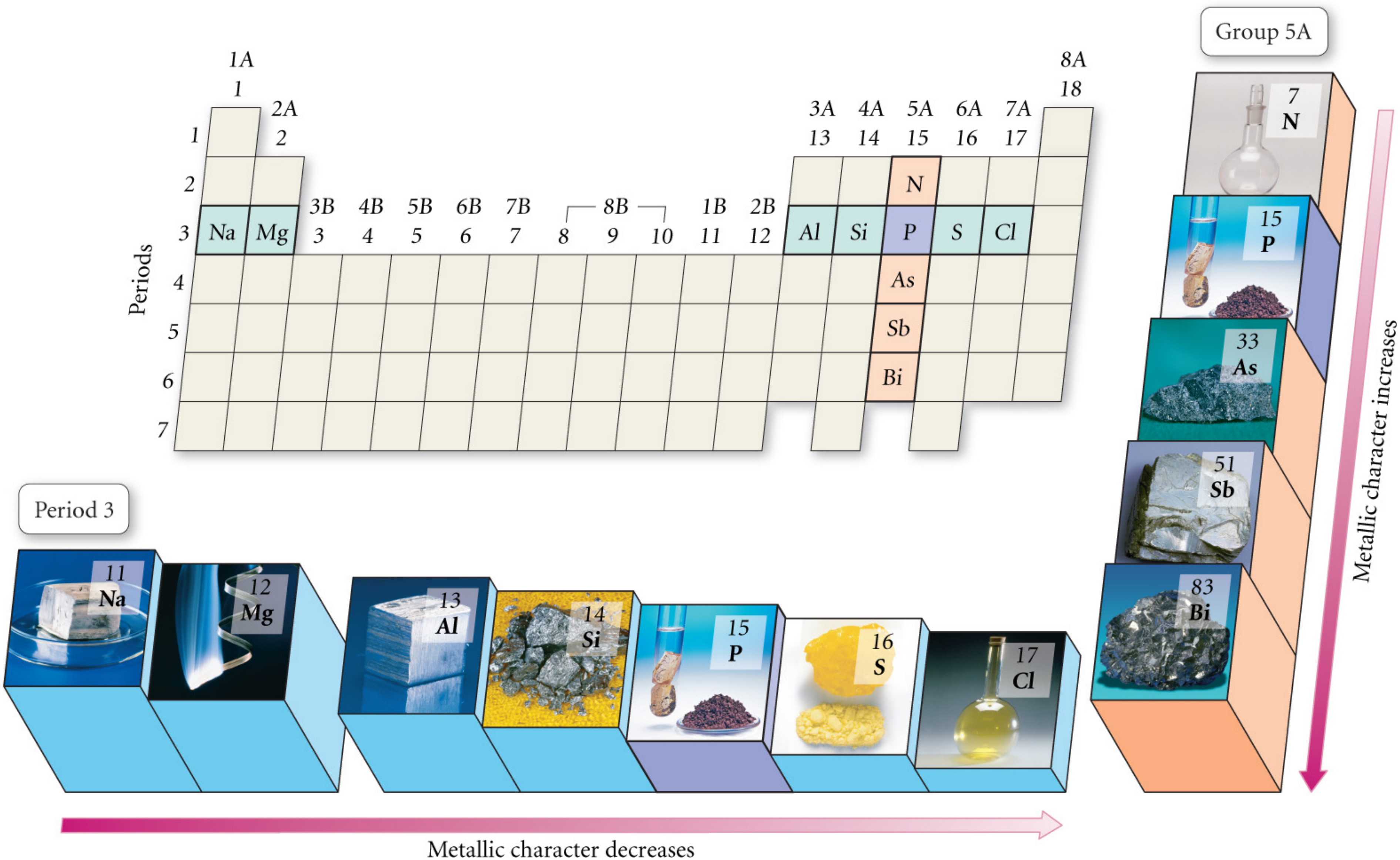
Chapter 11

Periodic Properties

- Metallic Character

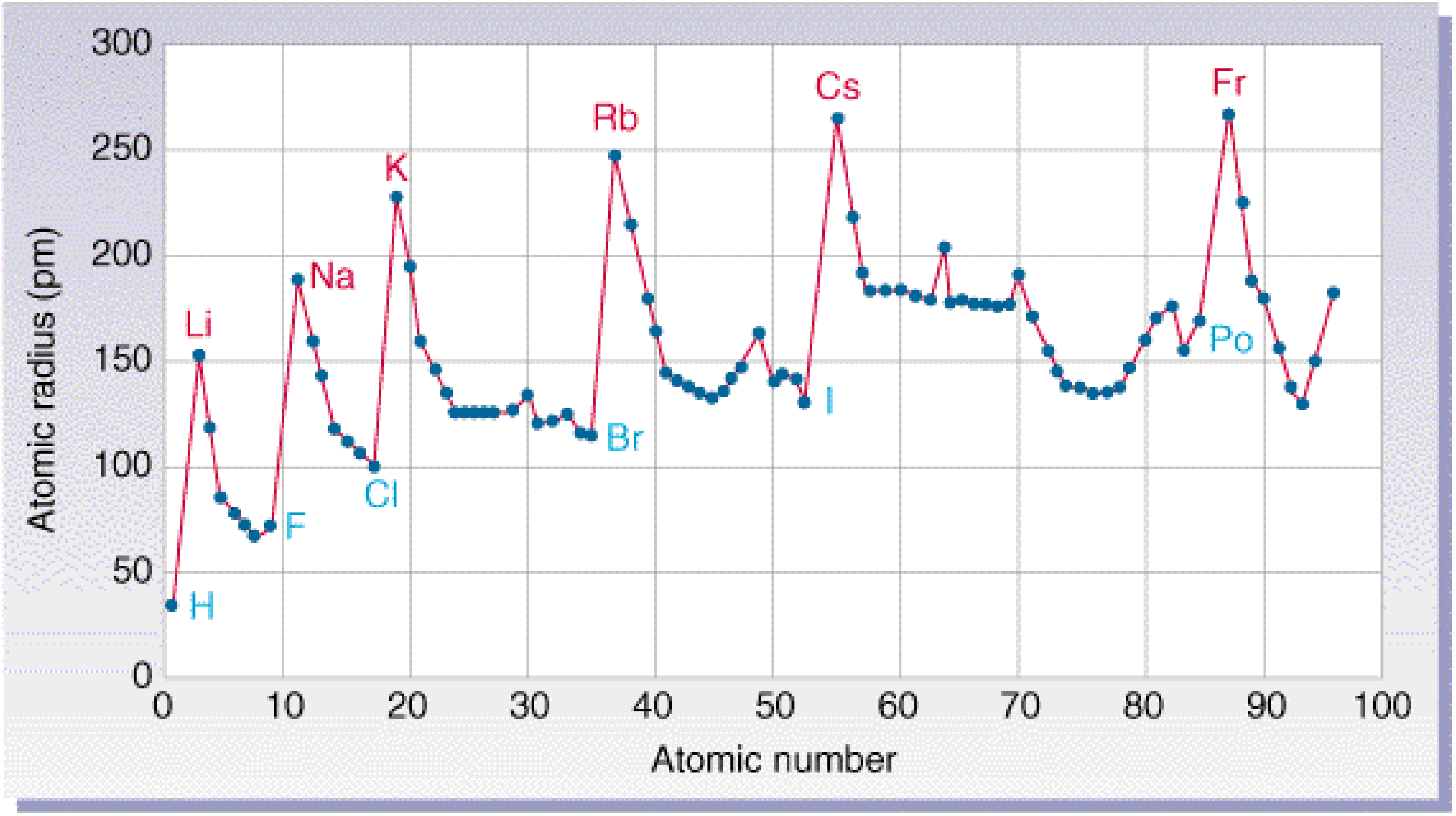


Trends in Metallic Character II

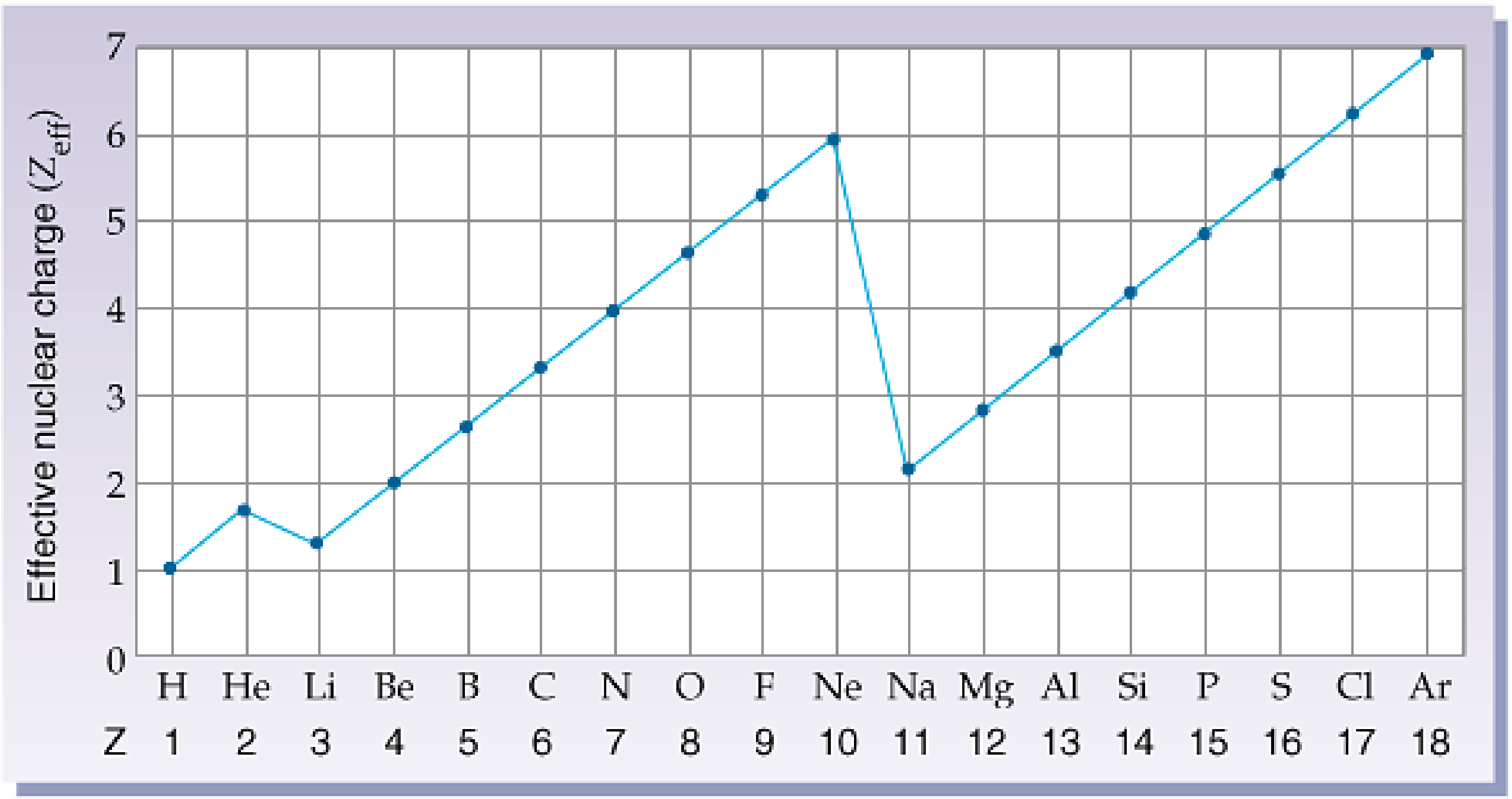


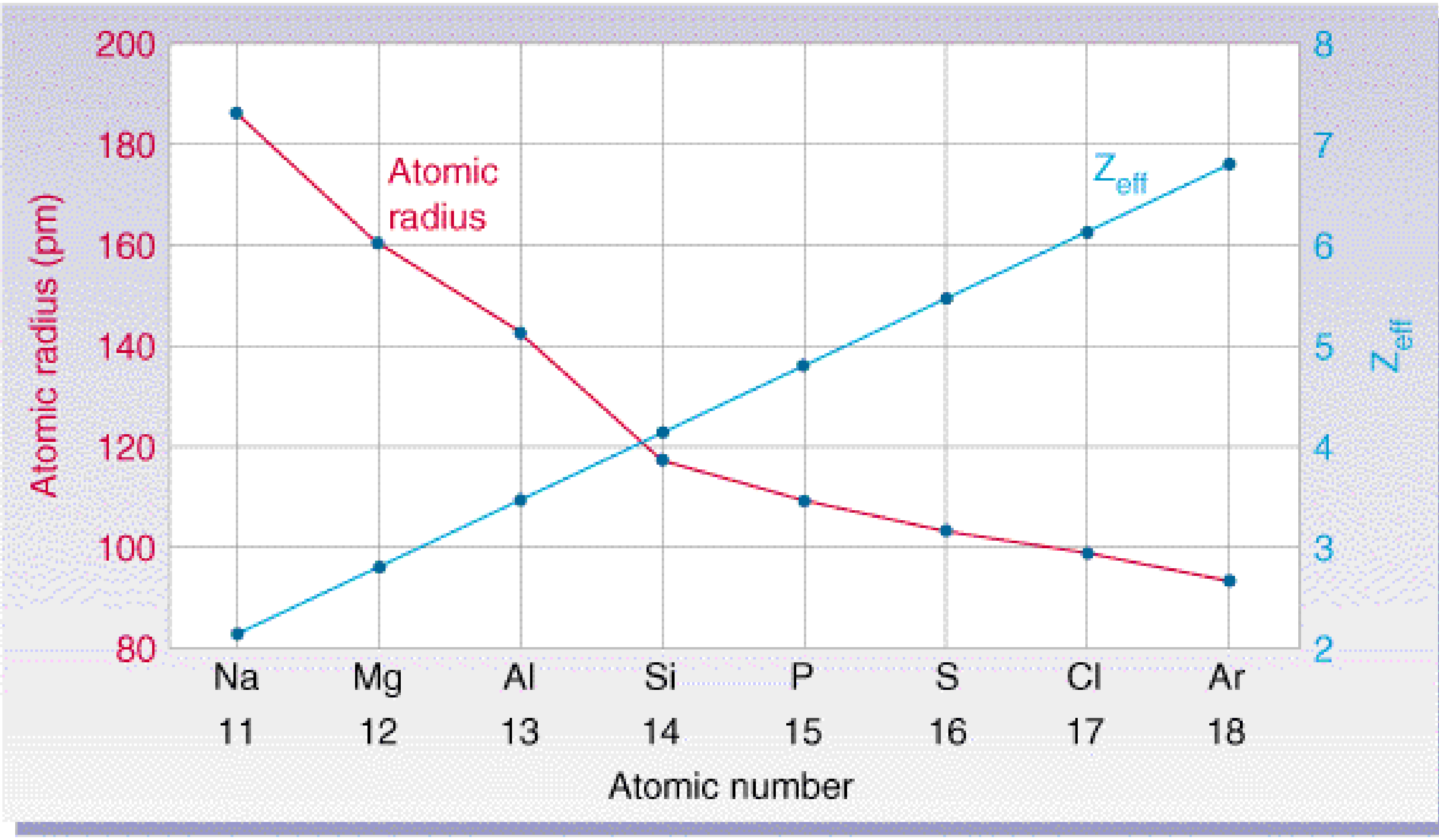
Periodic Properties

- Atomic Size – determined by how far the outermost electrons are from the nucleus



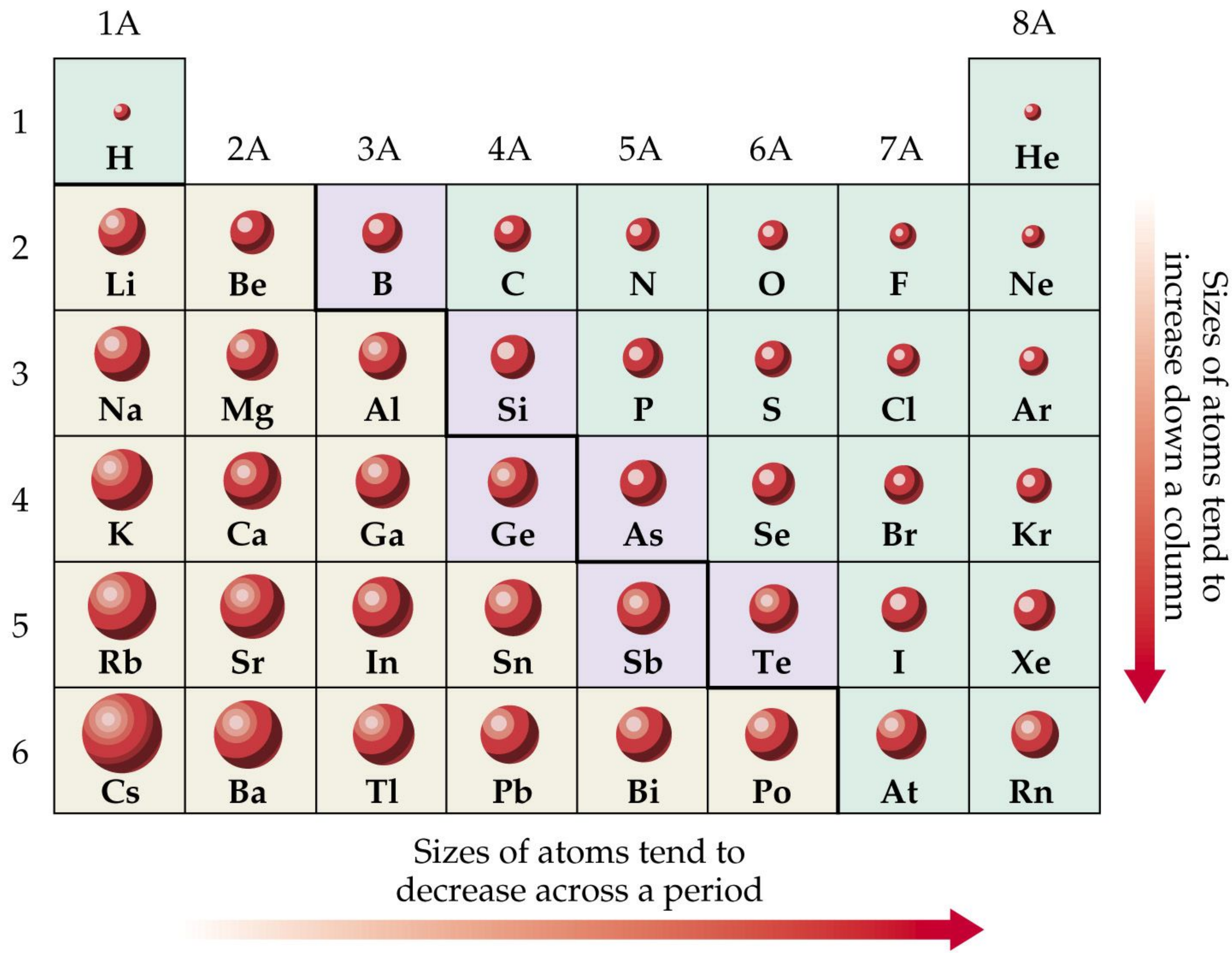
Effective nuclear charge





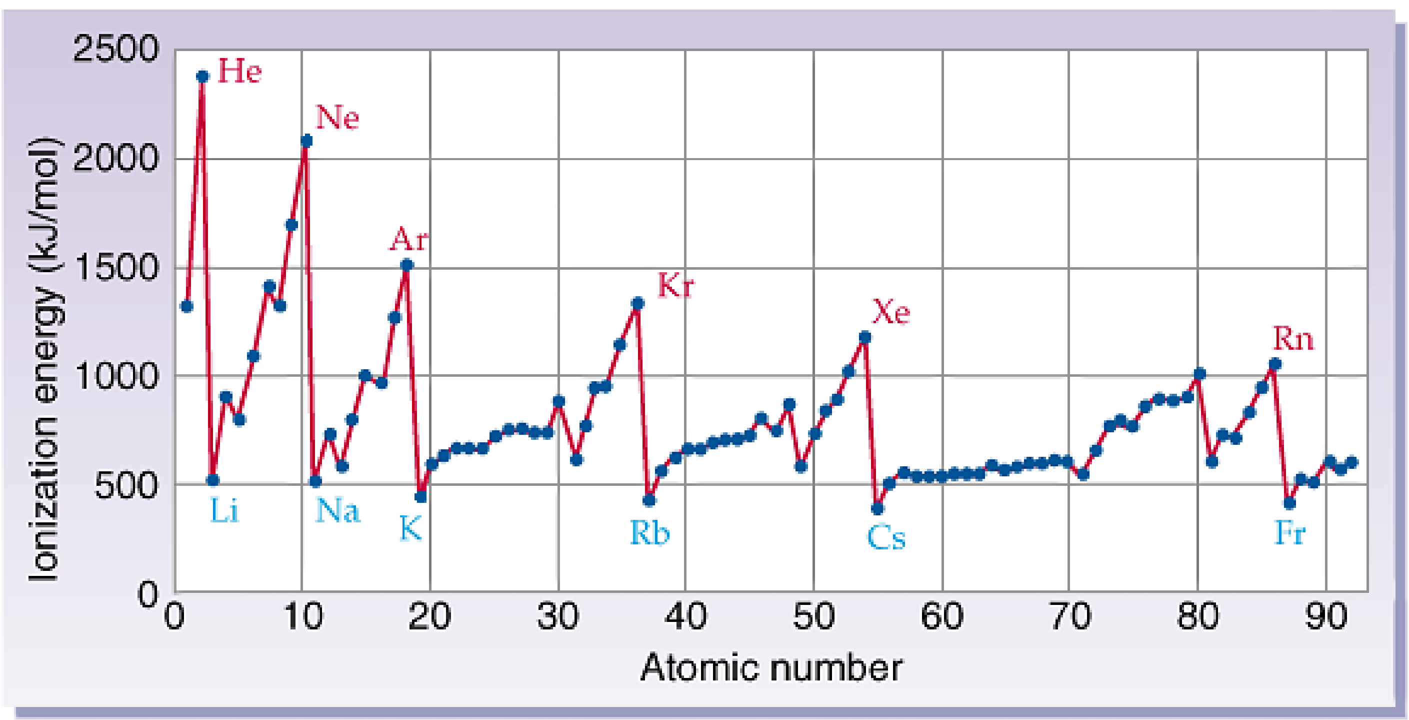
200
180
160
140
120
100
80
Atomic radius (pm)
8
7
6
5
4
3
2
Z_{eff}
Atomic radius
Z_{eff}
Na
Mg
Al
Si
P
S
Cl
Ar
11
12
13
14
15
16
17
18
Atomic number

Relative atomic sizes of the representative elements



Periodic Properties

- Ionization Energy - The amount of energy required to remove the outermost electron form an isolated neutral atom in the gaseous state.

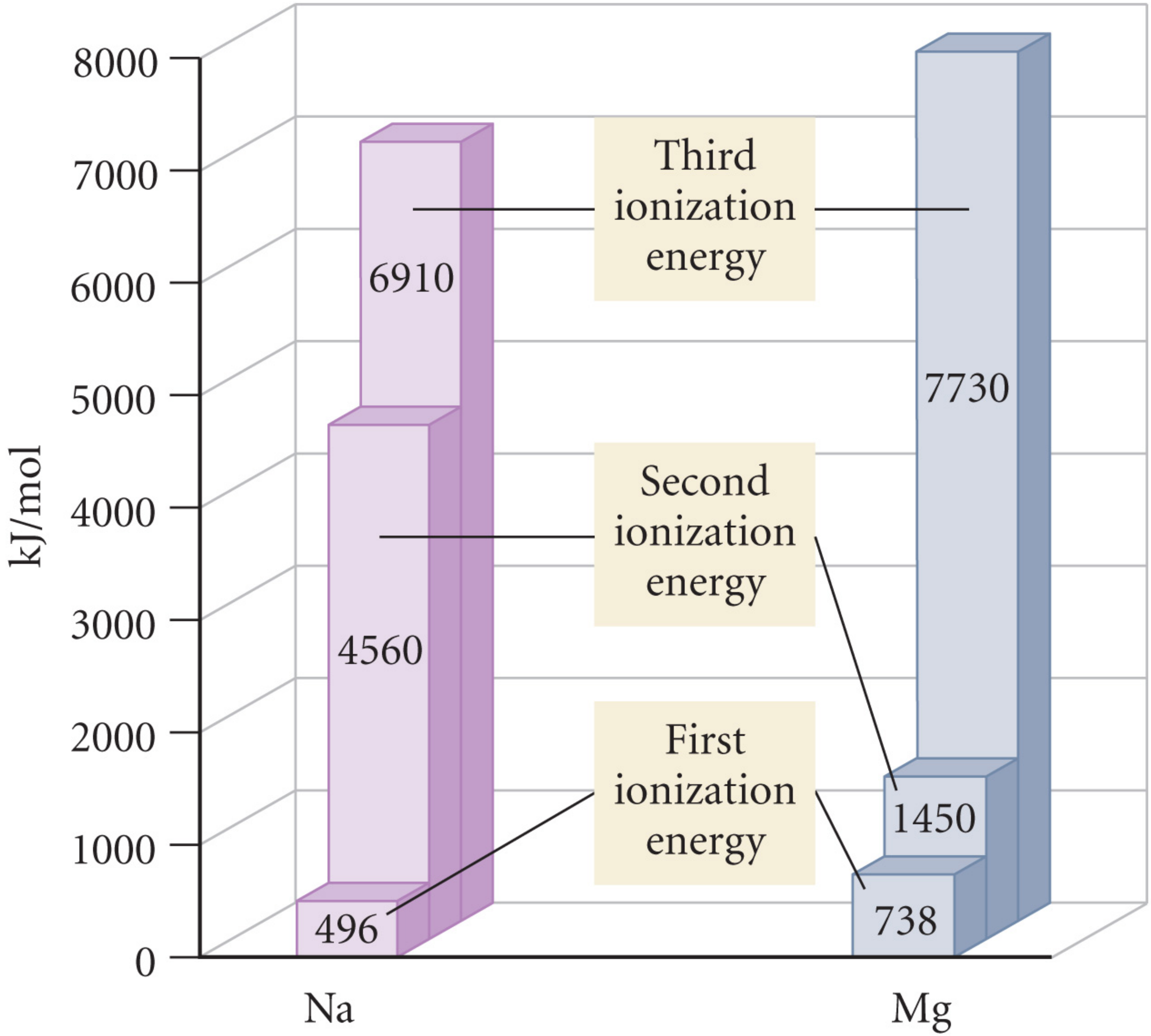


Ionization energy increases

Ionization energy decreases

Periods	1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1	1 **H**																	2 **He**
2	3 **Li**	4 **Be**											5 **B**	6 **C**	7 **N**	8 **O**	9 **F**	10 **Ne**
3	11 **Na**	12 **Mg**											13 **Al**	14 **Si**	15 **P**	16 **S**	17 **Cl**	18 **Ar**
4	19 **K**	20 **Ca**	21 **Sc**	22 **Ti**	23 **V**	24 **Cr**	25 **Mn**	26 **Fe**	27 **Co**	28 **Ni**	29 **Cu**	30 **Zn**	31 **Ga**	32 **Ge**	33 **As**	34 **Se**	35 **Br**	36 **Kr**
5	37 **Rb**	38 **Sr**	39 **Y**	40 **Zr**	41 **Nb**	42 **Mo**	43 **Tc**	44 **Ru**	45 **Rh**	46 **Pd**	47 **Ag**	48 **Cd**	49 **In**	50 **Sn**	51 **Sb**	52 **Te**	53 **I**	54 **Xe**
6	55 **Cs**	56 **Ba**	57 **La**	72 **Hf**	73 **Ta**	74 **W**	75 **Re**	76 **Os**	77 **Ir**	78 **Pt**	79 **Au**	80 **Hg**	81 **Tl**	82 **Pb**	83 **Bi**	84 **Po**	85 **At**	86 **Rn**
7	87 **Fr**	88 **Ra**	89 **Ac**	104 **Rf**	105 **Db**	106 **Sg**	107 **Bh**	108 **Hs**	109 **Mt**	110	111	112		114		116		

Lanthanides	58 **Ce**	59 **Pr**	60 **Nd**	61 **Pm**	62 **Sm**	63 **Eu**	64 **Gd**	65 **Tb**	66 **Dy**	67 **Ho**	68 **Er**	69 **Tm**	70 **Yb**	71 **Lu**
Actinides	90 **Th**	91 **Pa**	92 **U**	93 **Np**	94 **Pu**	95 **Am**	96 **Cm**	97 **Bk**	98 **Cf**	99 **Es**	100 **Fm**	101 **Md**	102 **No**	103 **Lr**


8000
7000
6000
5000
4000
3000
2000
1000
0
kJ/mol
Third ionization energy
Second ionization energy
First ionization energy
6910
4560
496
7730
1450
738
Na
Mg

Higher ionization energies

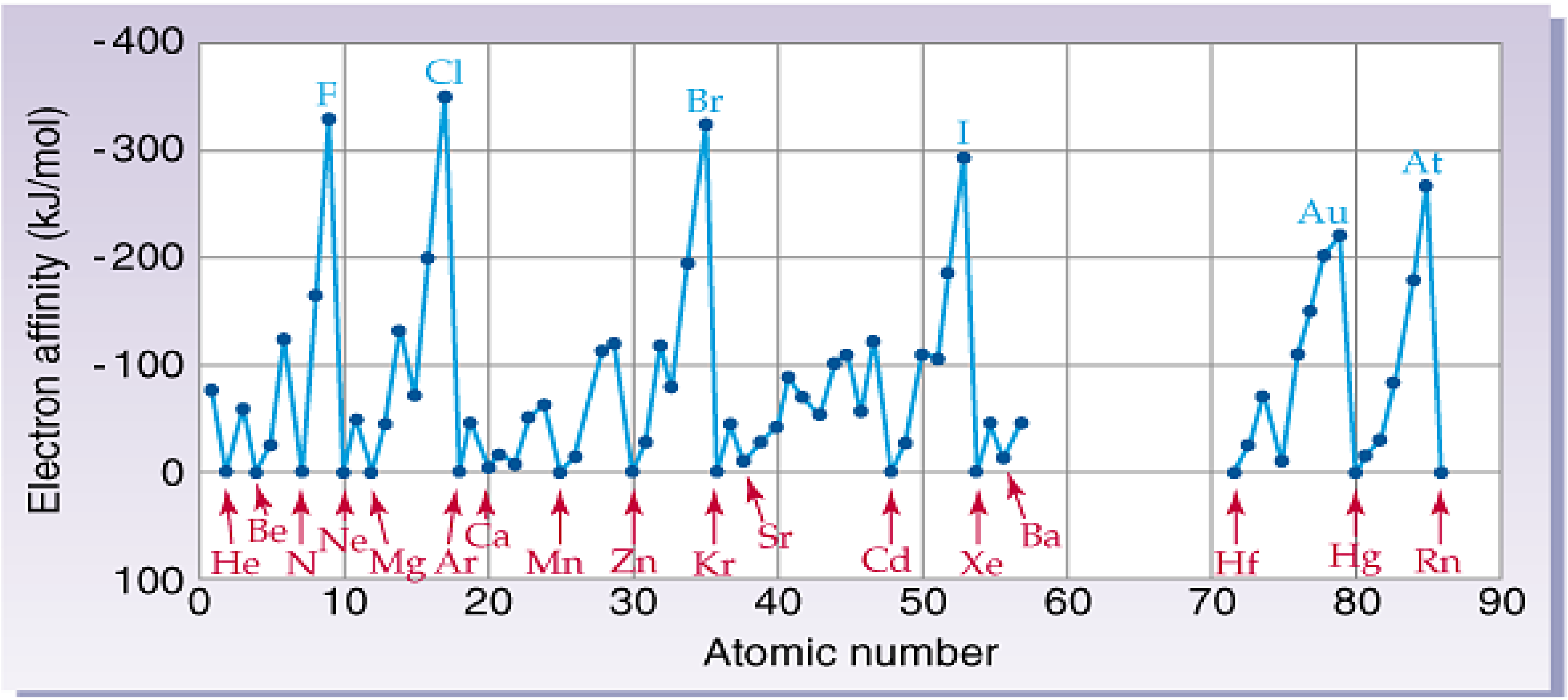
- Ionization energy always increases as you pull off more electrons.
- Ionization energies take a huge leap when we try to remove an electron from a new inner shell.

TABLE 6.1 Successive Ionization Energies (kJ/mol) for Third-Row Elements

E_i Number	Na	Mg	Al	Si	P	S	Cl	Ar
E_{i1}	496	738	578	787	1,012	1,000	1,251	1,520
E_{i2}	4,562	1,451	1,817	1,577	1,903	2,251	2,297	2,665
E_{i3}	6,912	7,733	2,745	3,231	2,912	3,361	3,822	3,931
E_{i4}	9,543	10,540	11,575	4,356	4,956	4,564	5,158	5,770
E_{i5}	13,353	13,630	14,830	16,091	6,273	7,013	6,540	7,238
E_{i6}	16,610	17,995	18,376	19,784	22,233	8,495	9,458	8,781
E_{i7}	20,114	21,703	23,293	23,783	25,397	27,106	11,020	11,995

Electron Affinity

- The energy change that occurs when an electron is added to an atom (or ion) in the gaseous state. Frequently costs nothing but actually yields energy therefore EA's are usually negative.



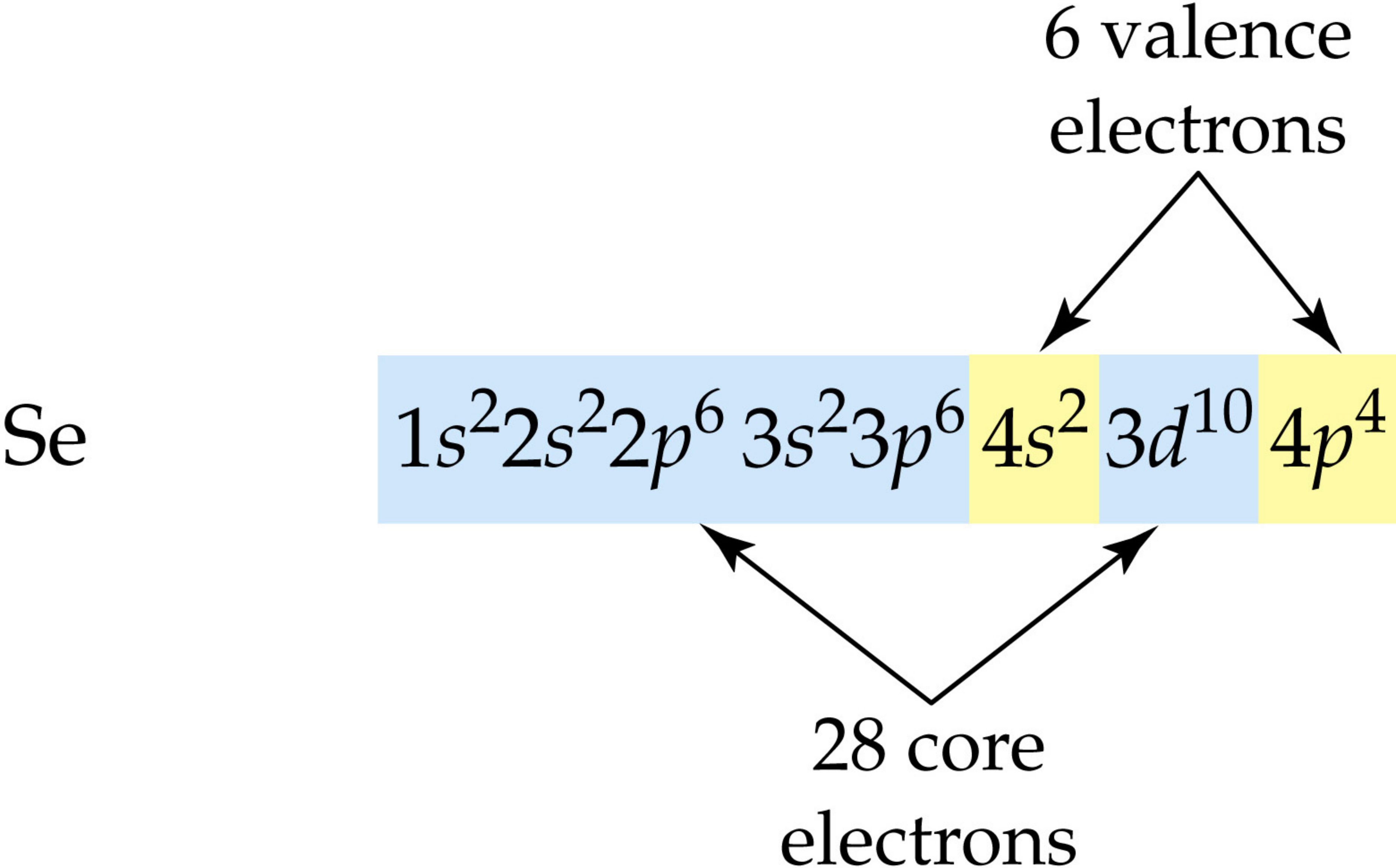
Valence Electrons

- The electrons that occupy the outermost s and p orbitals of an atom

Si $1s^2 2s^2 2p^6 \, 3s^2 3p^2$

Core electrons

Valence electrons


6 valence electrons
Se
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$
28 core electrons

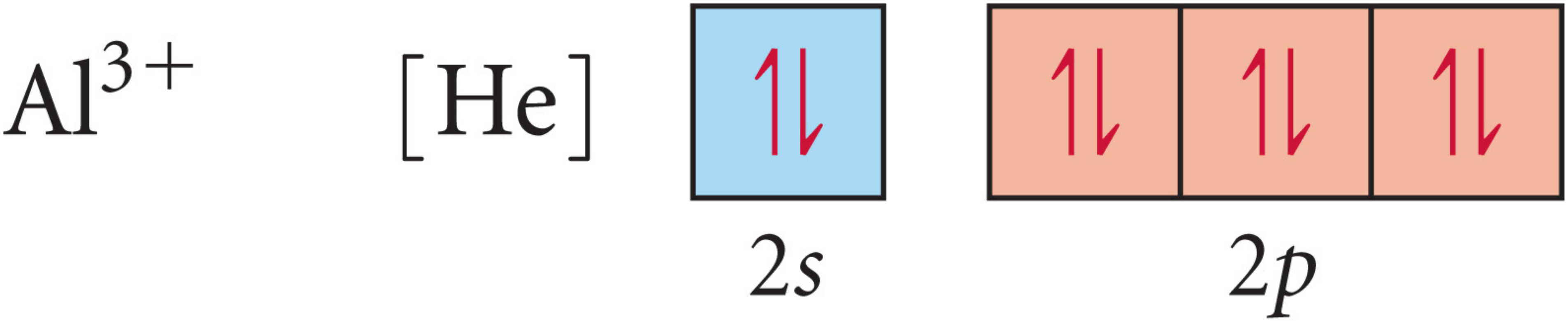
Alkali metals	Alkaline earth metals	Noble gases	Halogens
3 **Li** $2s^1$	4 **Be** $2s^2$	2 **He** $1s^2$	9 **F** $2s^2 2p^5$
11 **Na** $3s^1$	12 **Mg** $3s^2$	10 **Ne** $2s^2 2p^6$	17 **Cl** $3s^2 3p^5$
19 **K** $4s^1$	20 **Ca** $4s^2$	18 **Ar** $3s^2 3p^6$	35 **Br** $4s^2 4p^5$
37 **Rb** $5s^1$	38 **Sr** $5s^2$	36 **Kr** $4s^2 4p^6$	53 **I** $5s^2 5p^5$
55 **Cs** $6s^1$	56 **Ba** $6s^2$	54 **Xe** $5s^2 5p^6$	85 **At** $6s^2 6p^5$
87 **Fr** $7s^1$	88 **Ra** $7s^2$	86 **Rn** $6s^2 6p^6$	

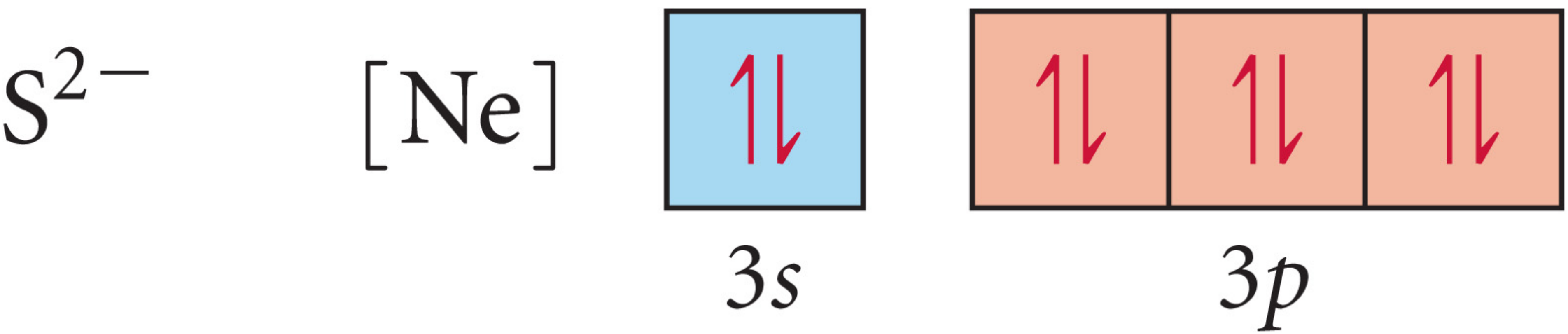
Electronic Configuration of the Ions

- Cations - Electrons are removed from the highest energy occupied orbital

- Anions - Electrons are added to the lowest energy unoccupied orbital

- For transition metals -- The highest ns electrons are removed first (even though they are not the last added)


Al^{3+}
[He]
2s
2p


S^{2-}
[Ne]
3s
3p

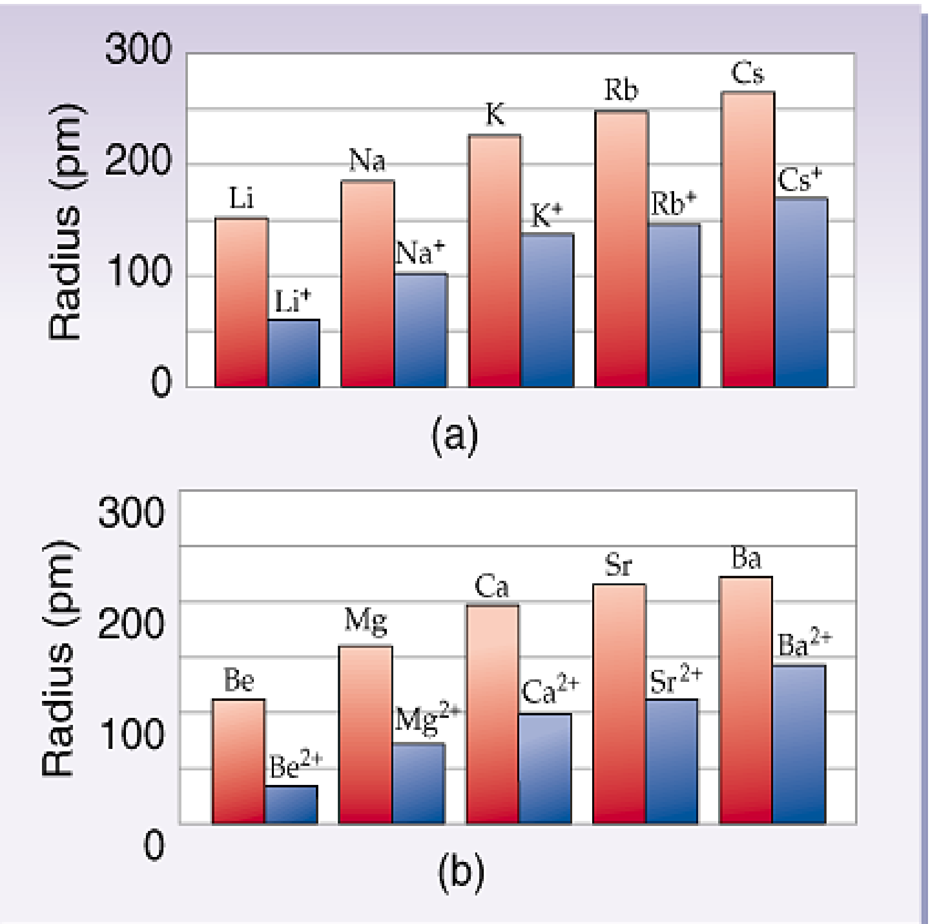
Isoelectronic

- Isoelectronic species have the same electron configuration.
- Atoms tend to gain or lose electrons to become isoelectronic with noble gases

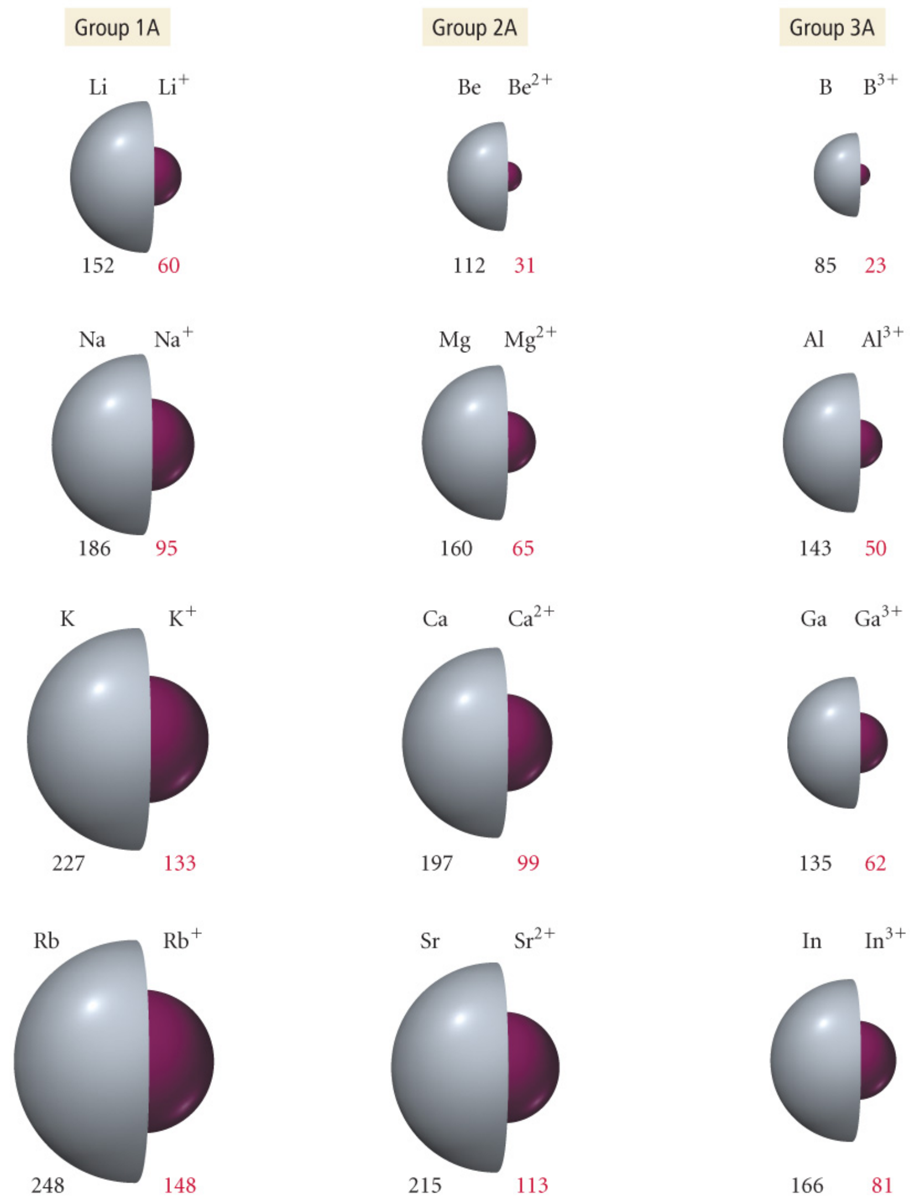
- Ne $1s^2\ 2s^2\ 2p^6$
- Na^{+1} $1s^2\ 2s^2\ 2p^6$
- F^{-1} $1s^2\ 2s^2\ 2p^6$

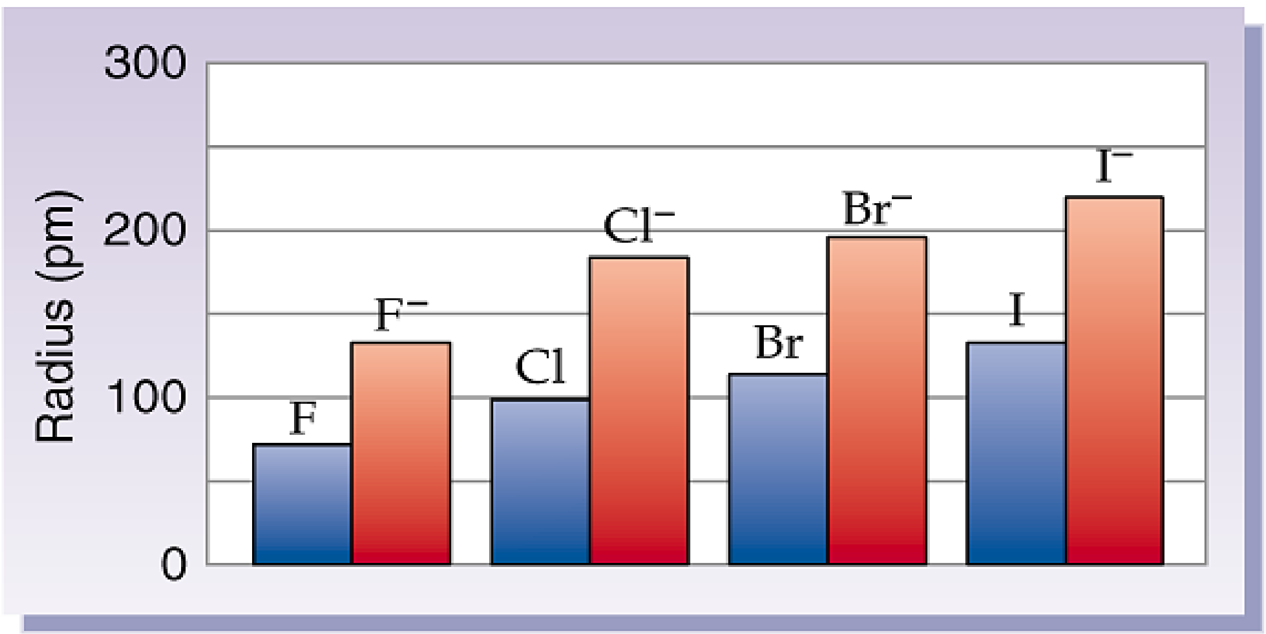
Ionic Radii

- Cations -- radius decreases due to an increase in Zeffective
- Anions -- radius increases due to crowding of more electrons into a shell

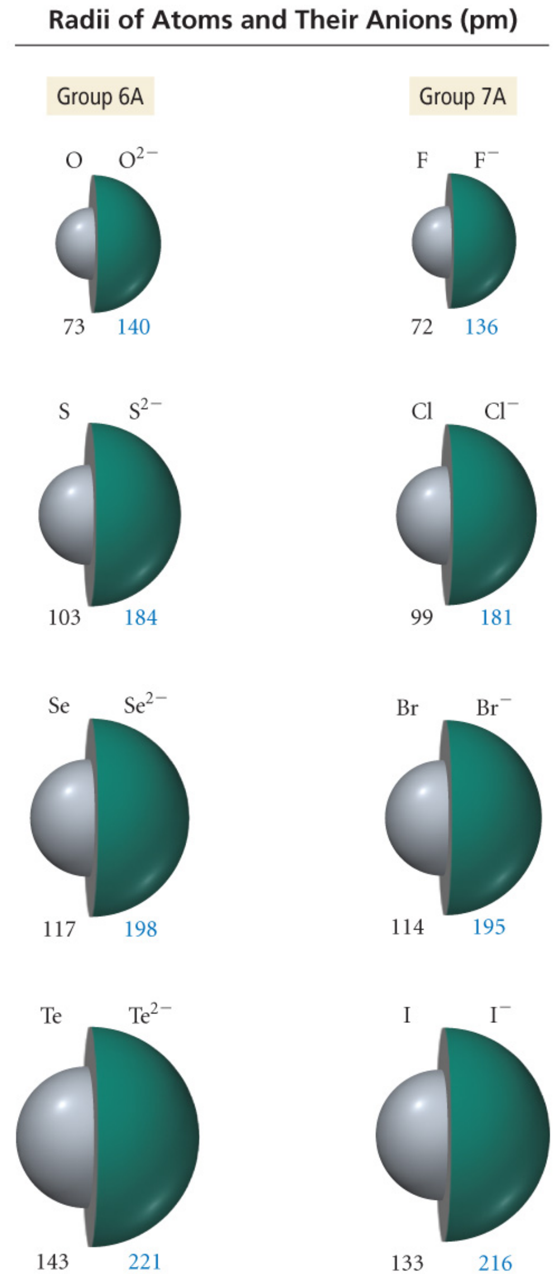


Radii of Atoms and Their Cations (pm)





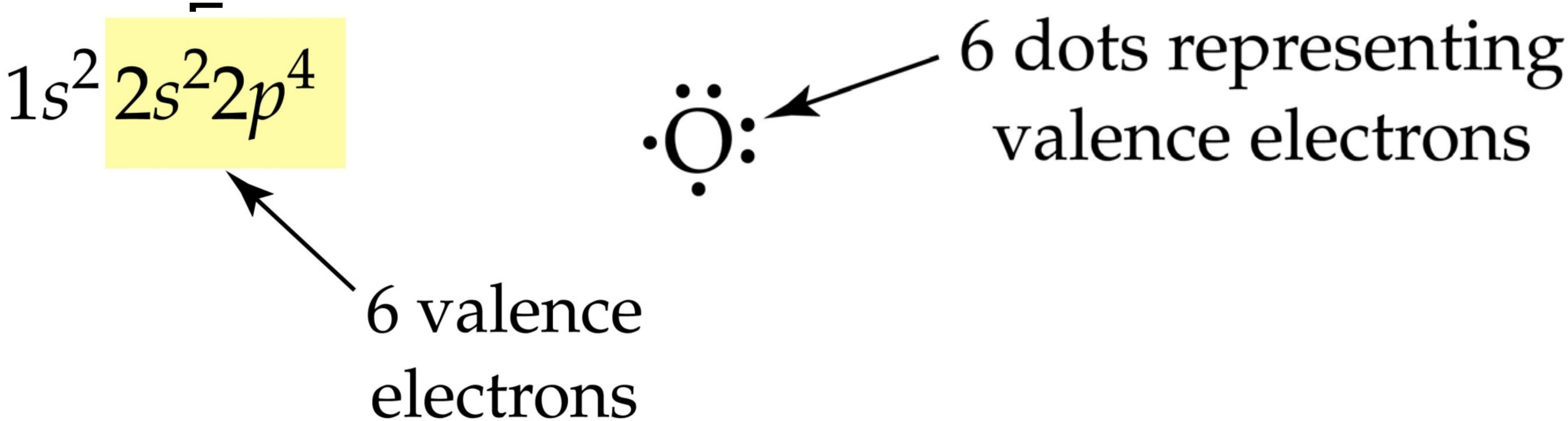
Radius (pm)
300
200
100
0
F
F⁻
Cl
Cl⁻
Br
Br⁻
I
I⁻



Radii of Atoms and Their Anions (pm)
Group 6A
Group 7A
O O^{2-}
73 140
F F^-
72 136
S S^{2-}
103 184
Cl Cl^-
99 181
Se Se^{2-}
117 198
Br Br^-
114 195
Te Te^{2-}
143 221
I I^-
133 216

Lewis Electron Dot Structures

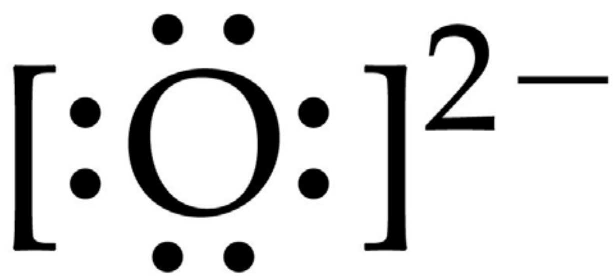
- For elements
  - Composed of elemental symbol + dots representing the outer shell or valence electrons



Lewis Electron Dot Structures

- For ions
  - Add or subtract dots for electrons gained or lost to form ion.
  - For O^{2-} –

$1s^2 2s^2 2p^6$



Bonding

- Atoms like to have a full outer shell and will gain lose or share electrons to achieve a full outer shell

- Representative elements gain, lose, or share electrons to have 8 electrons in their outer shell corresponding to full s and p orbitals.
  - This is the <u>octet rule.</u>

Two kinds of bonds

- **Ionic**
  - atoms gain or lose electrons to form octet
  - ions held together by electrostatic forces
- **Covalent**
  - atoms share electrons to form octet
  - atoms held together by shared electron covalent bonds

Ion Formation

- Cations
- $Na \rightarrow Na^{+} + 1e^{-}$
- $Mg \rightarrow Mg^{+2} + 2e^{-}$
-
- Anions
- $Cl + 1\ e^{-} \rightarrow Cl^{-1}$
- $S + 2\ e^{-} \rightarrow S^{-2}$

These elements tend to lose electrons to gain an octet

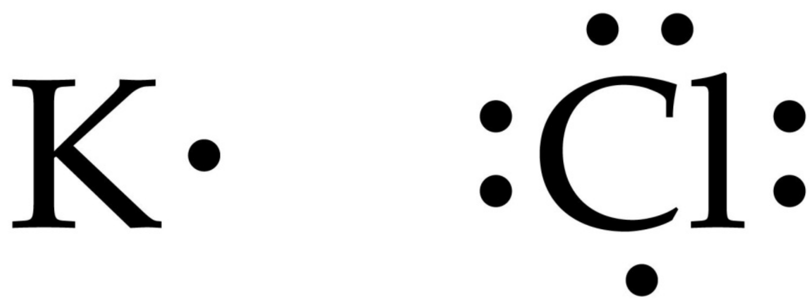
Note: losing an electron always costs energy -- but sometimes this lost of energy can be compensated by the strong electrostatic energy gained when a cation and an anion combined.

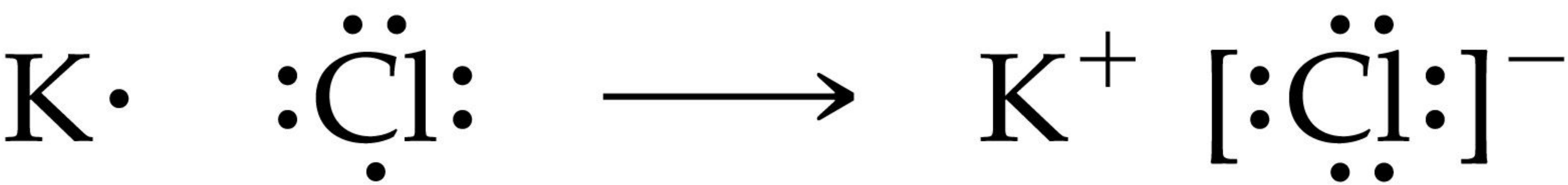
These elements tend to gain electrons to form an octet

Ionic Bonds

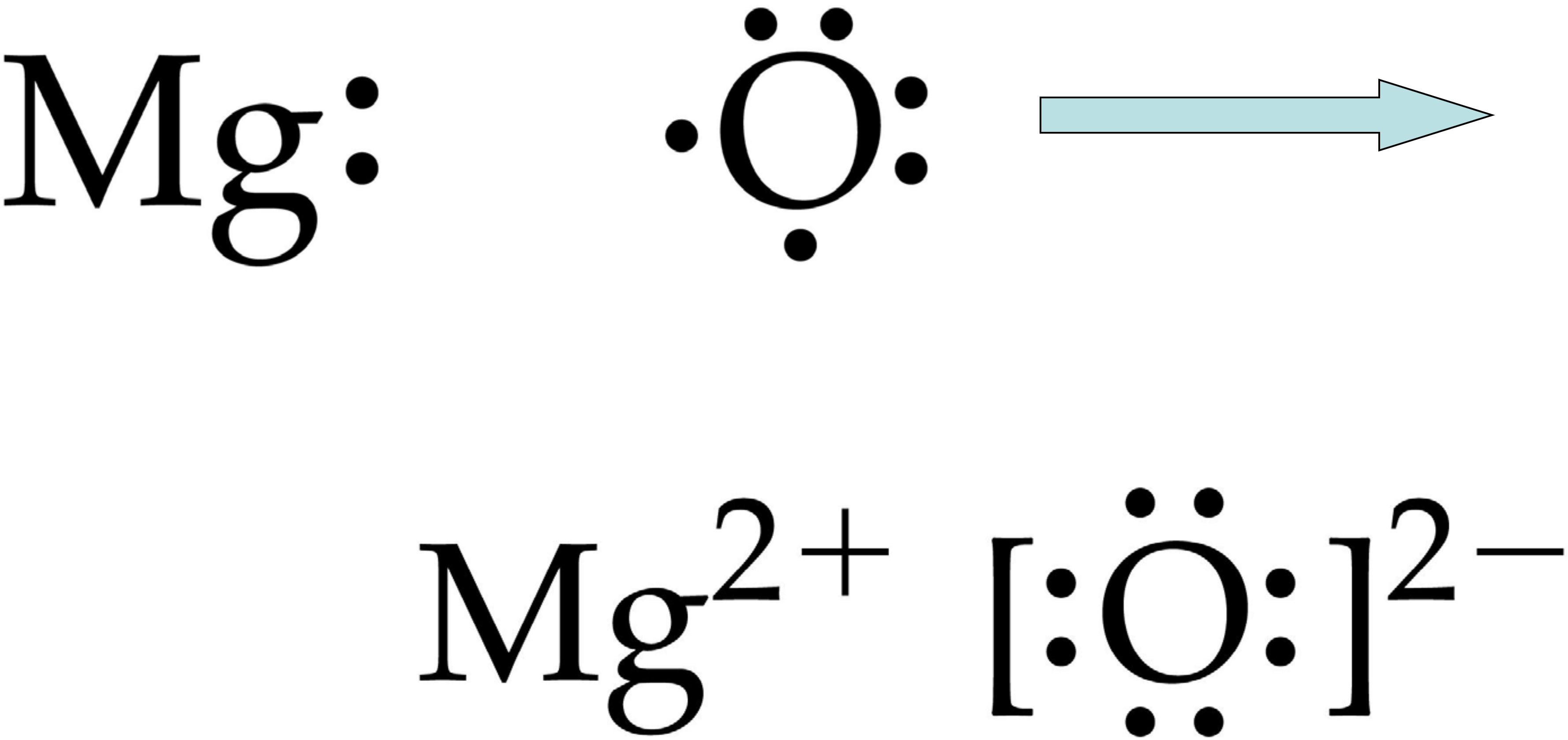
- Bonds formed by the interaction of ions and the strong electrostatic forces that hold them together.
- Ions group together in ratios which balance their positive and negative charges → results in a neutral crystal

What compound will be formed by the reaction of potassium and chlorine?

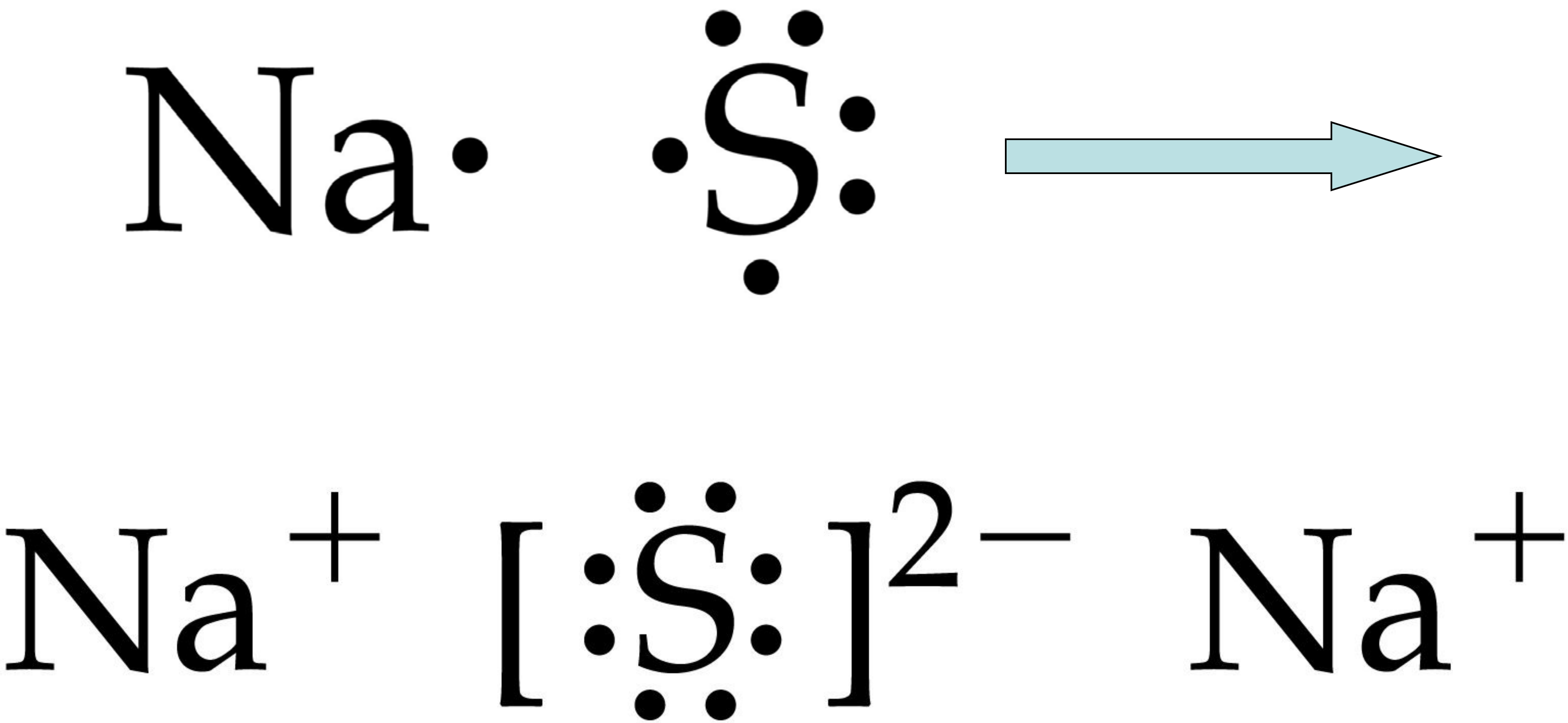




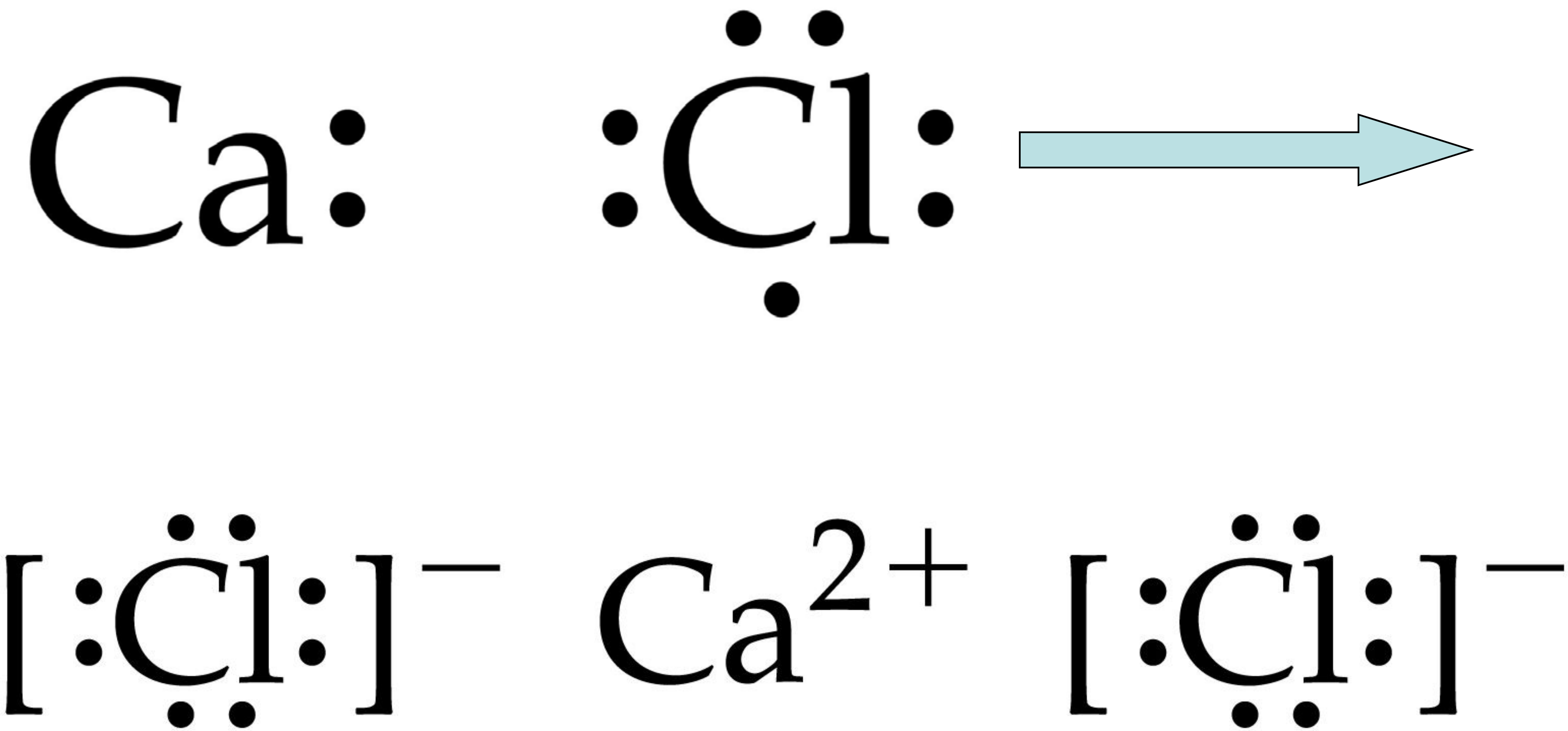
What compound will be formed by the reaction of magnesium and oxygen?



What compound will be formed by the reaction of sodium and sulfur?

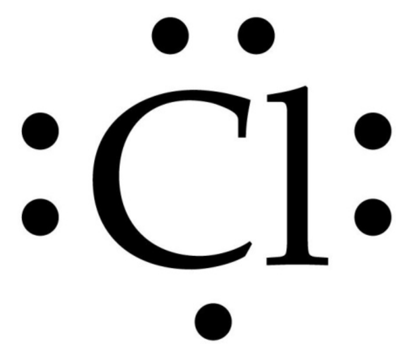


What compound will be formed by the reaction of sodium and sulfur?



Covalent Bonds

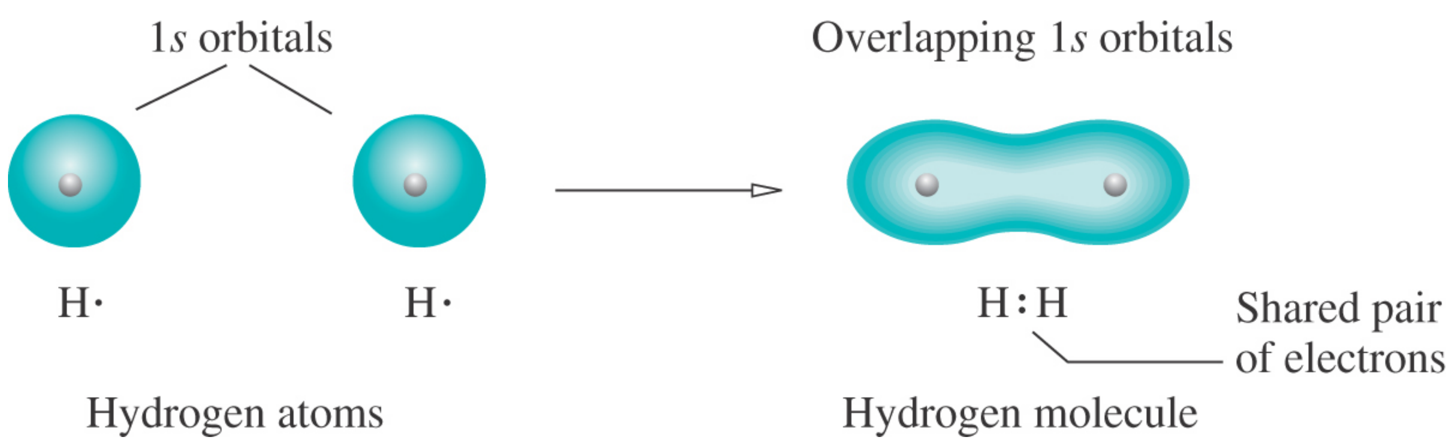
- Bonds where atoms share electrons to achieve an optimum number of electrons in their outer shells. Typically an octet.



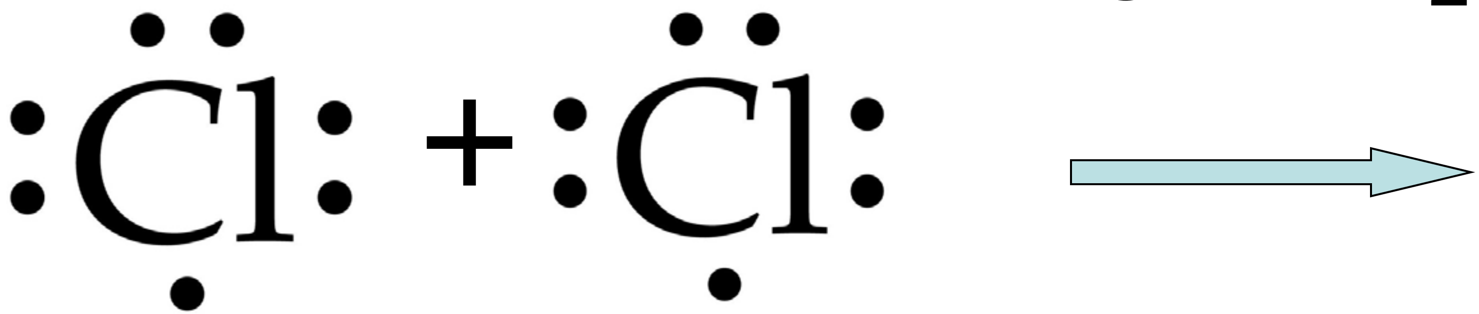
Show bonding in H_2

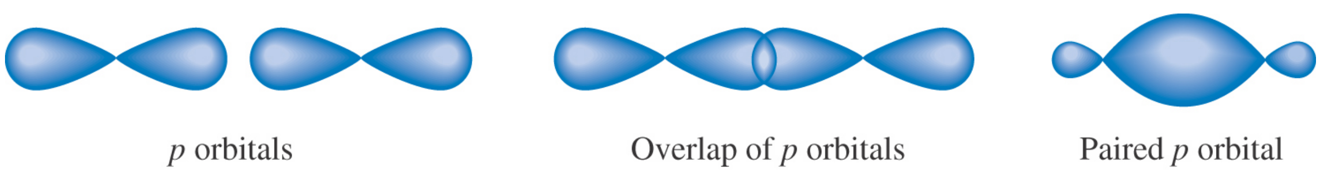
$$H\cdot + H\cdot \longrightarrow$$

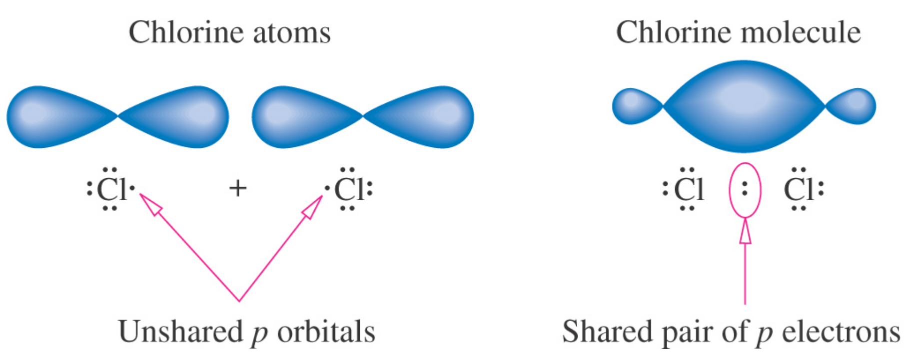
$$H{:}H \text{ or } H\text{—}H$$



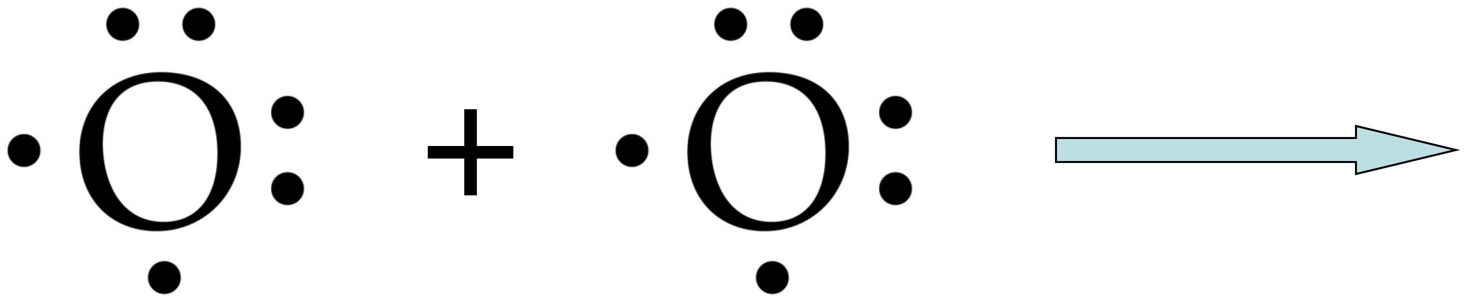
Show bonding in Cl_2

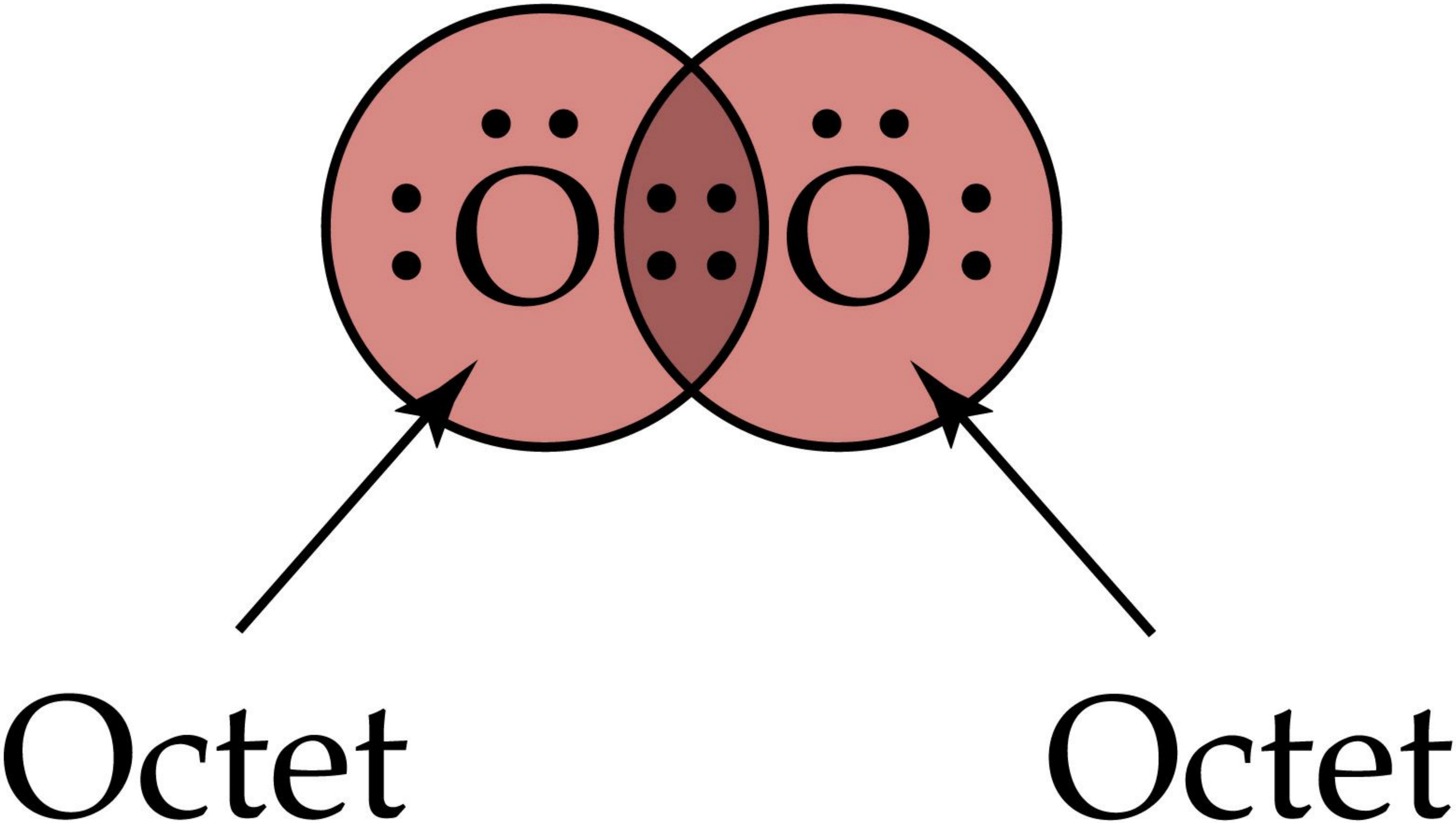






Show bonding in O_2



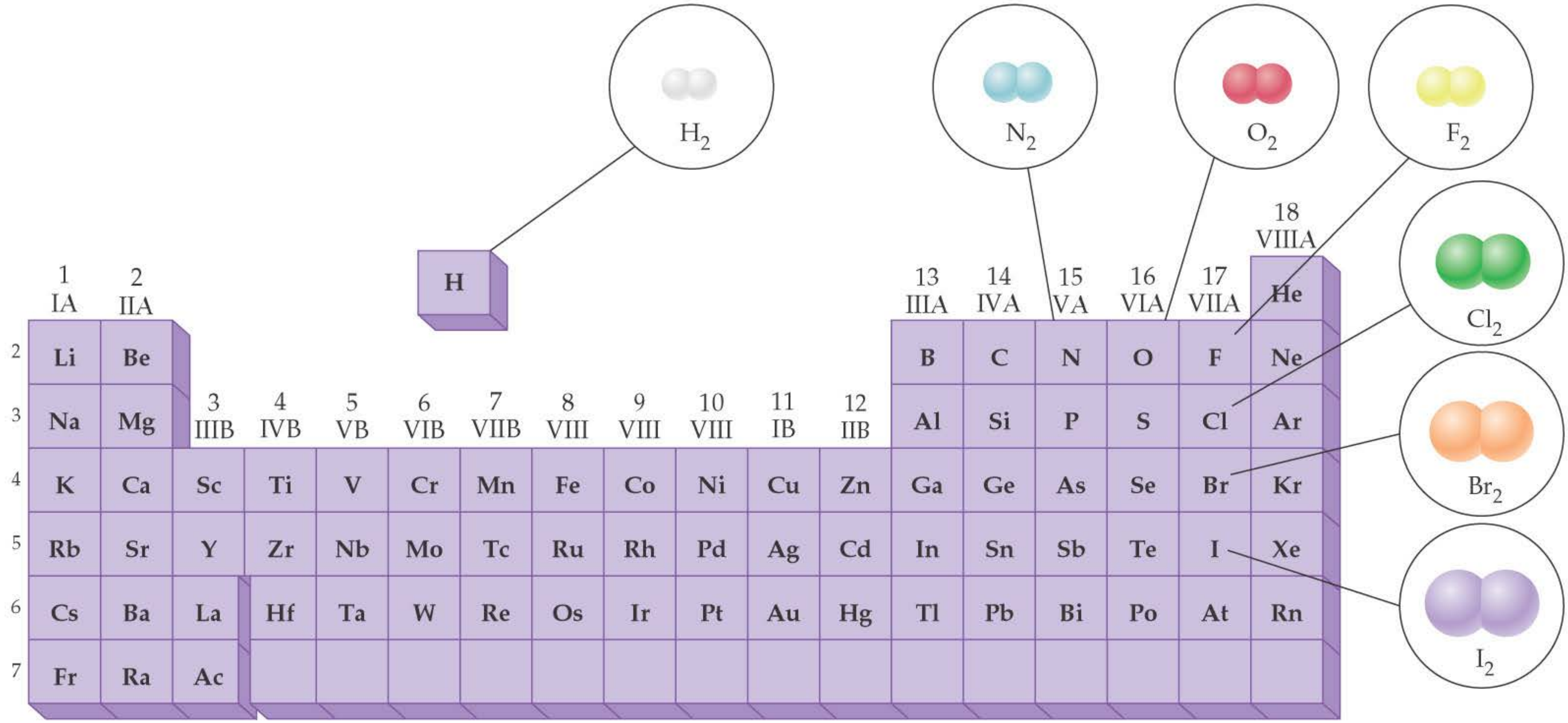


:Ö::Ö:
Octet
Octet

Show bonding in N_2

$$:\mathrm{N}:::\mathrm{N}: \quad \text{or} \quad :\mathrm{N}\equiv\mathrm{N}:$$

Diatomic Elements



Bond Strength

- A measure of the amount of energy it takes to break a bond.

Bond Length

- The length of a bond between two atoms is the distance separation the nuclei of the atoms.

Bonding in nonmetals

- Generally every unpaired electron in the Lewis Dot diagram of an element can form a bond.

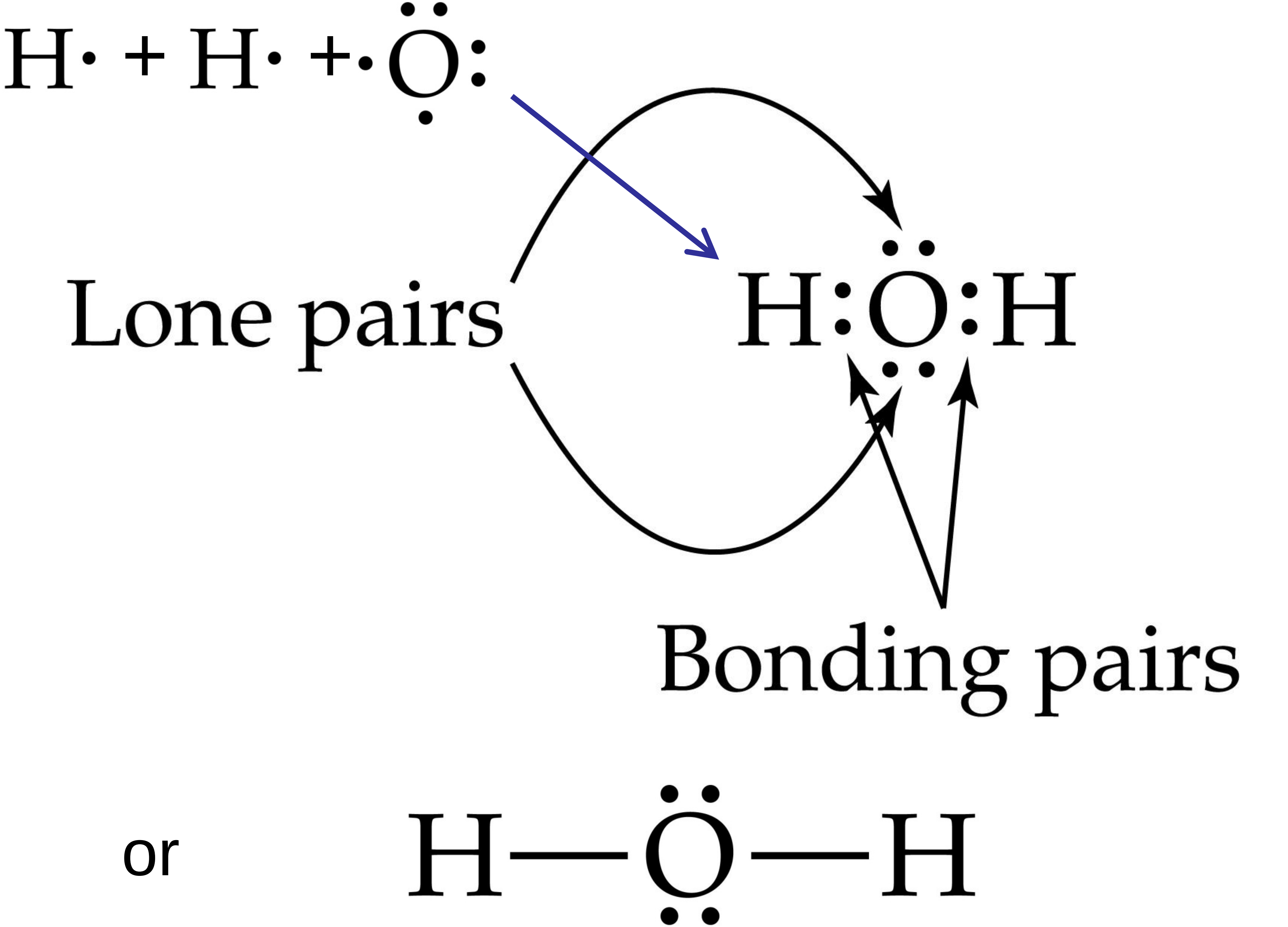
Bonding in nonmetals

- Generally every unpaired electron in the Lewis Dot diagram of an element can form a bond.

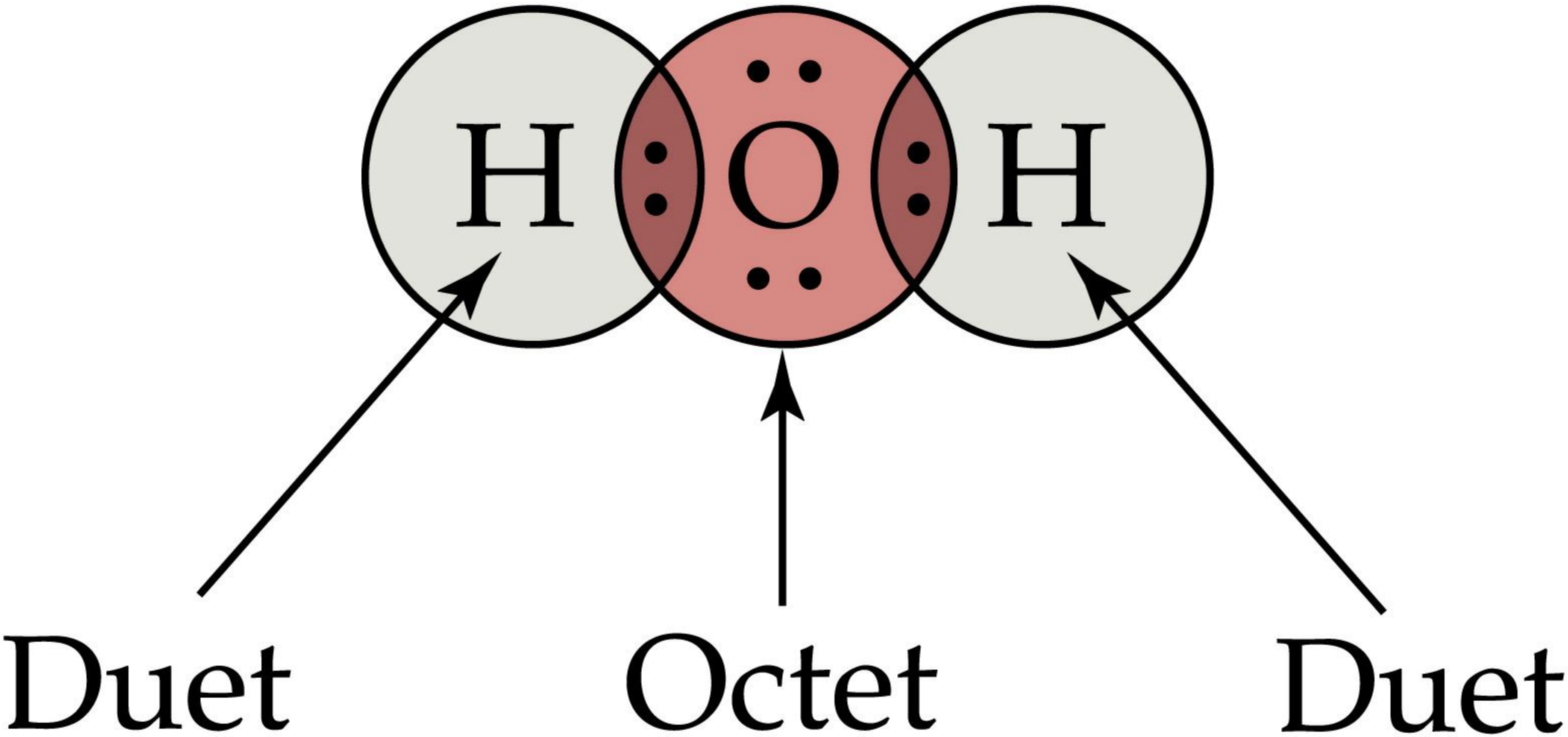
· B ·	· C ·	· N ·	: O ·	: F :
3(4) bonds	4 bonds	3 bonds	2 bonds	1 bond

Lewis Electron Dot Structures

- Bonding electrons pairs – electron pairs involved in bonds
- Lone electron pairs – electron pairs that do not participate in bonding
- Bond order = number of bonds



Lone pairs
Bonding pairs
or


H
O
H
Duet
Octet
Duet

Writing Lewis Dot Structures

- Decide which atoms are bonded together - draw a skeleton structure
- Count the total number of valence electrons available.
- Find the number of electrons needed to give an octet around all atoms -- (remember H needs 2, all else need 8).

Writing Lewis Dot Structures

- Determine number of electrons short.
- Number of bonds needed = number of electrons short/2.
- Distribute bonds -- (1st hook atoms together and then add double bonds where appropriate).
- Calculate number of electrons used in bonds.

Writing Lewis Dot Structures

- Calculate electrons remaining.
- Distribute remaining electrons to give all atoms an octet.
- Done!!

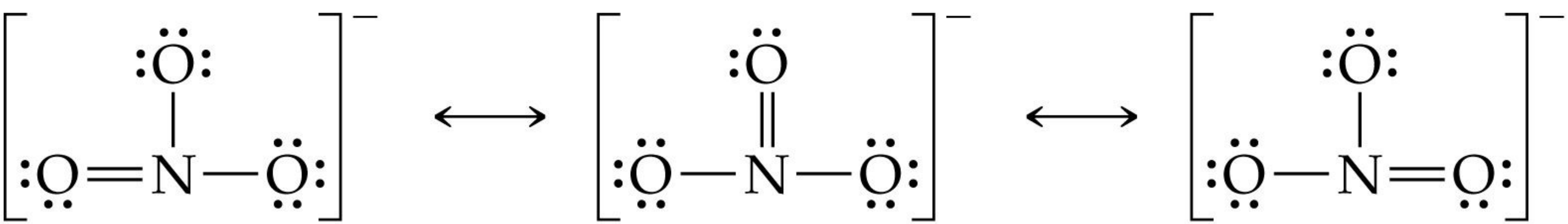
Lewis Structures of ions

- for anions add the extra electrons to the number available
- for cations subtract the lost electrons from the number available

Resonance

- In some Lewis structures, the multiple bonds can be written in several equivalent locations. All structures have the exact same energy. Which is the correct Lewis structure??

- Answer : None alone are correct – the true molecule is a hybrid of the possible structures. The electrons are **delocalized**.

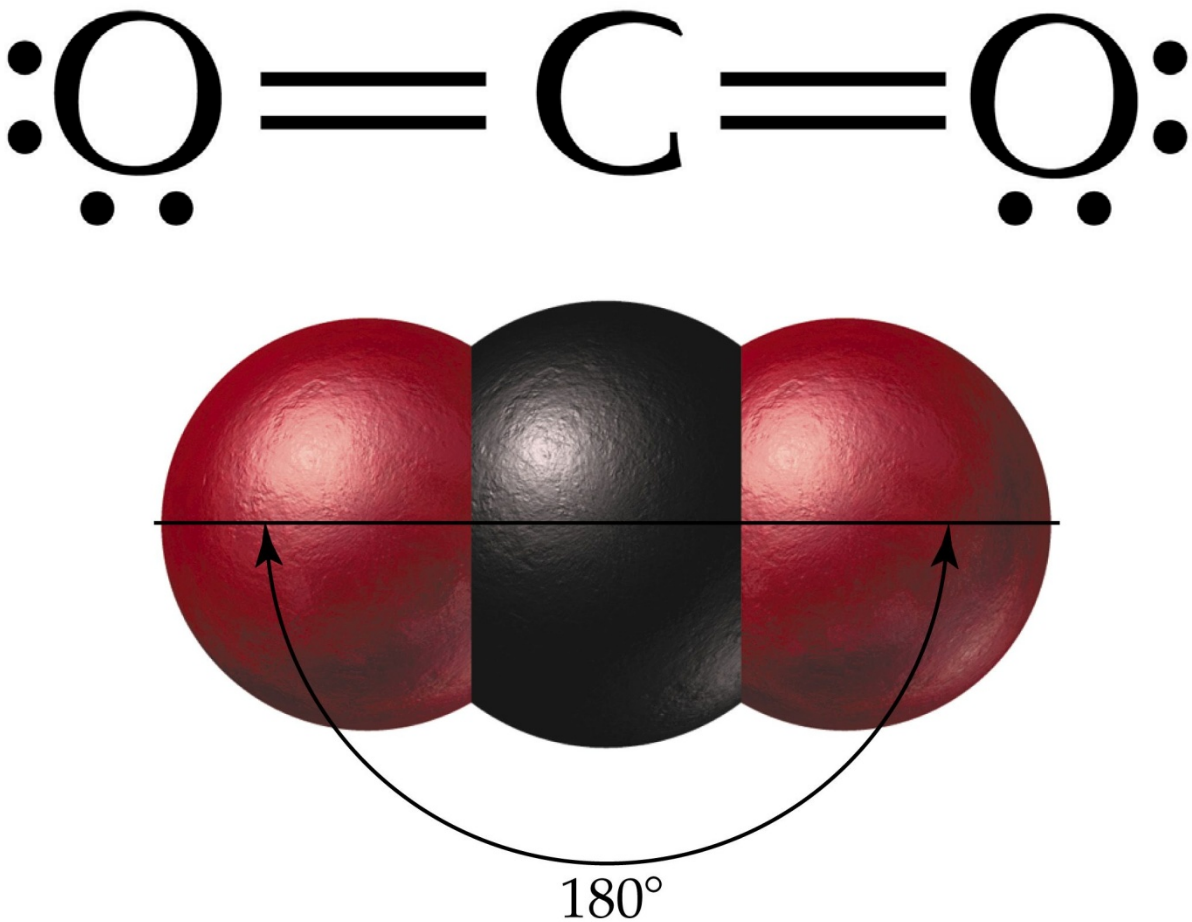


Predicting Shapes of Molecules

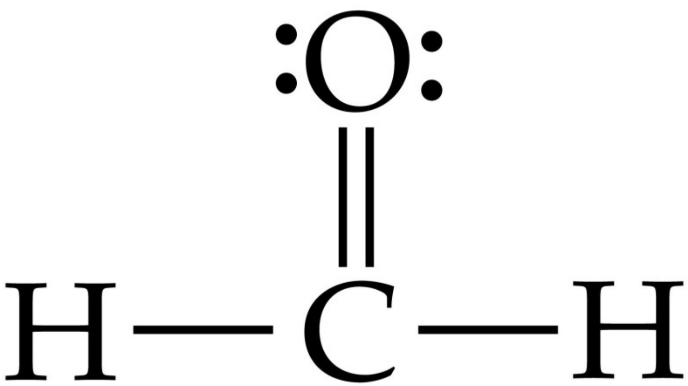
- To predict the shapes of molecules we look at the things (sigma bonds or lone pairs of electrons) surrounding them and put them as far from each other as possible.

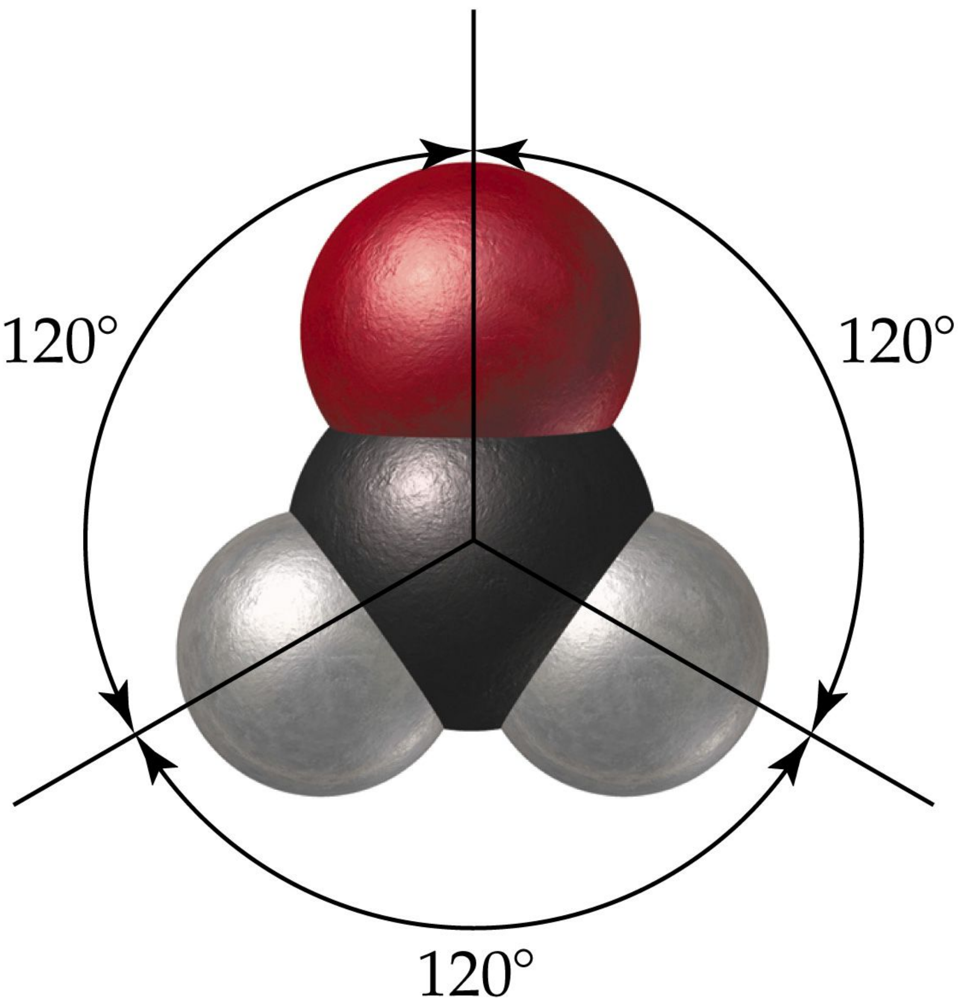
- Valence Shell Electron Pair Repulsion (VSEPR) Theory

2 atoms attached to central atom

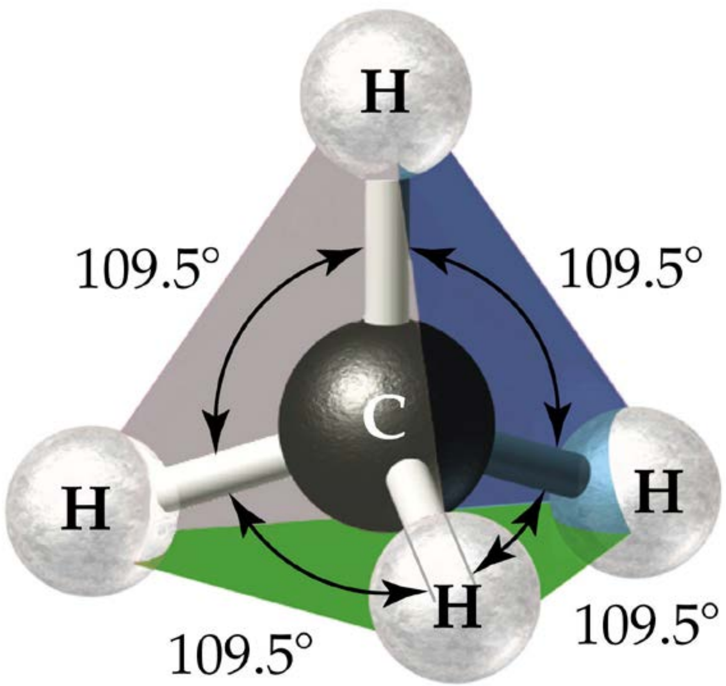


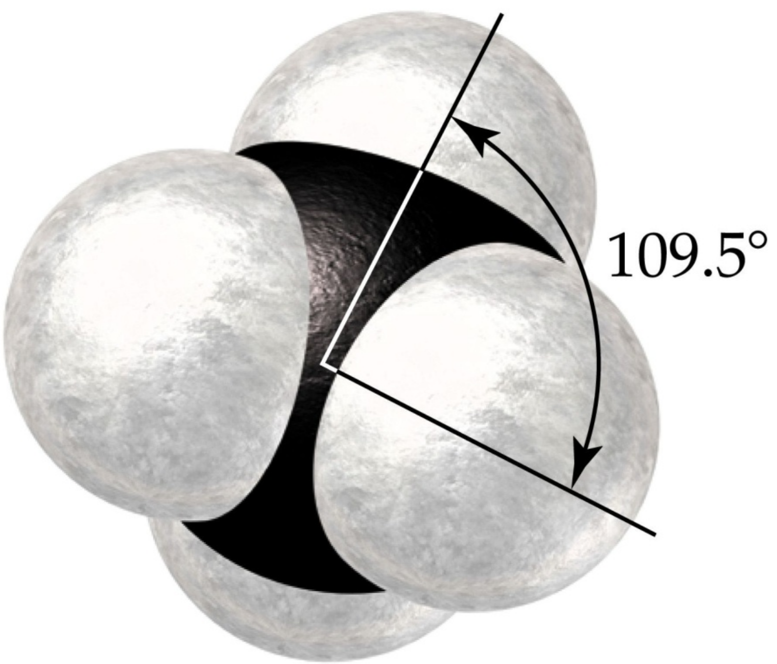
3 atoms attached to central atom



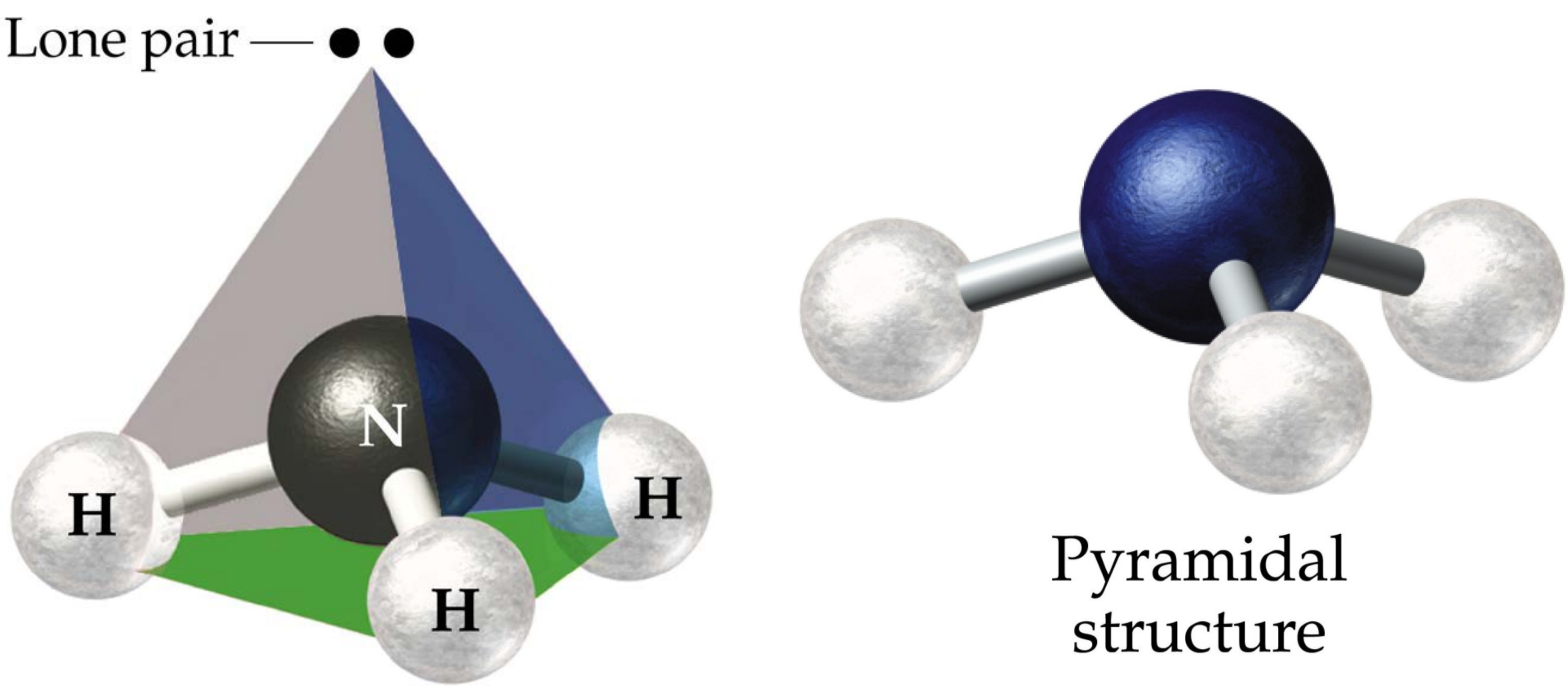


4 atoms attached to central atom

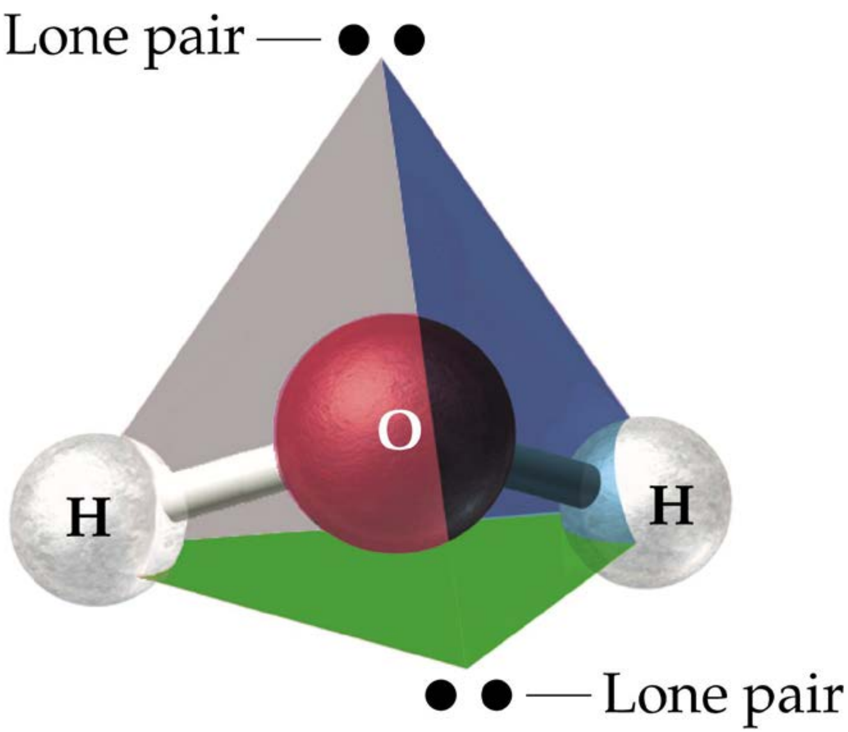




3 atoms + lone pair attached to central atom



2 atoms + 2 lone pairs attached to central atom



Bent structure

Table 11.6 Arrangement of Electron Pairs and Molecular Structure

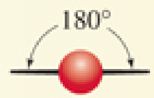
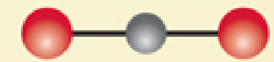
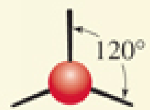
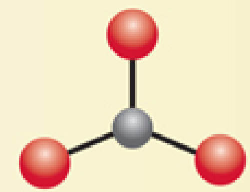
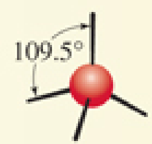
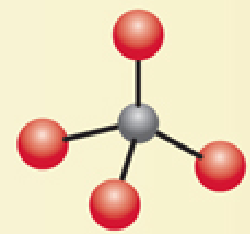
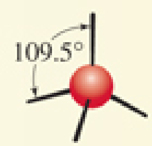
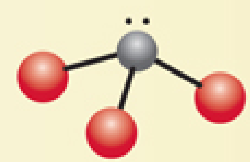
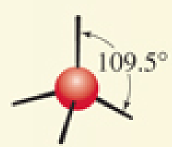
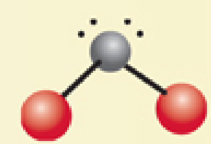
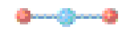
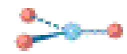
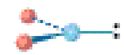
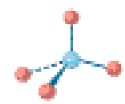
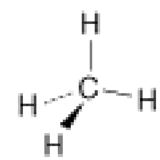
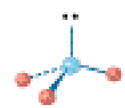
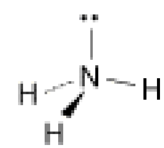
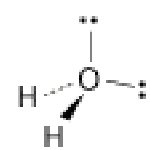
Number of electron pairs	Electron pair arrangement	Ball-and-stick model	Bonds	Molecular structure	Molecular structure model
2	Linear		2	Linear	
3	Trigonal planar		3	Trigonal planar	
4	Tetrahedral		4	Tetrahedral	
4	Tetrahedral		3	Trigonal pyramidal	
4	Tetrahedral		2	Bent	

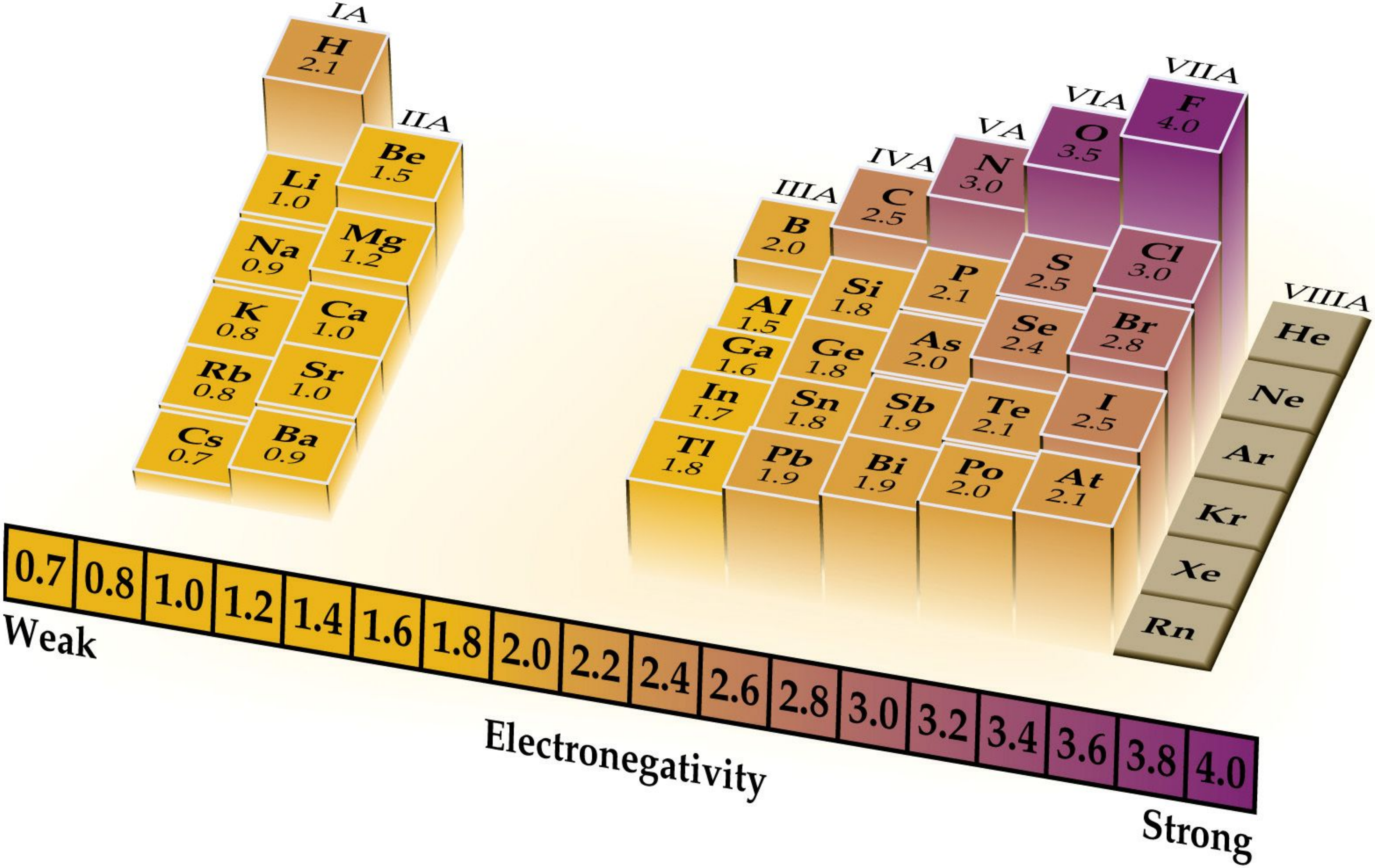
TABLE 7.4 Molecular Geometry Around Atoms with 2, 3, 4, 5, and 6 Charge Clouds

Number of Bonds	Number of Lone Pairs	Number of Charge Clouds	Molecular Geometry	Example
2	0	2	 Linear	Cl—Be—Cl
3	0	3	 Trigonal planar	
2	1	3	 Bent	
4	0	4	 Tetrahedral	
3	1	4	 Trigonal pyramidal	
2	2	4	 Bent	

Electronegativity

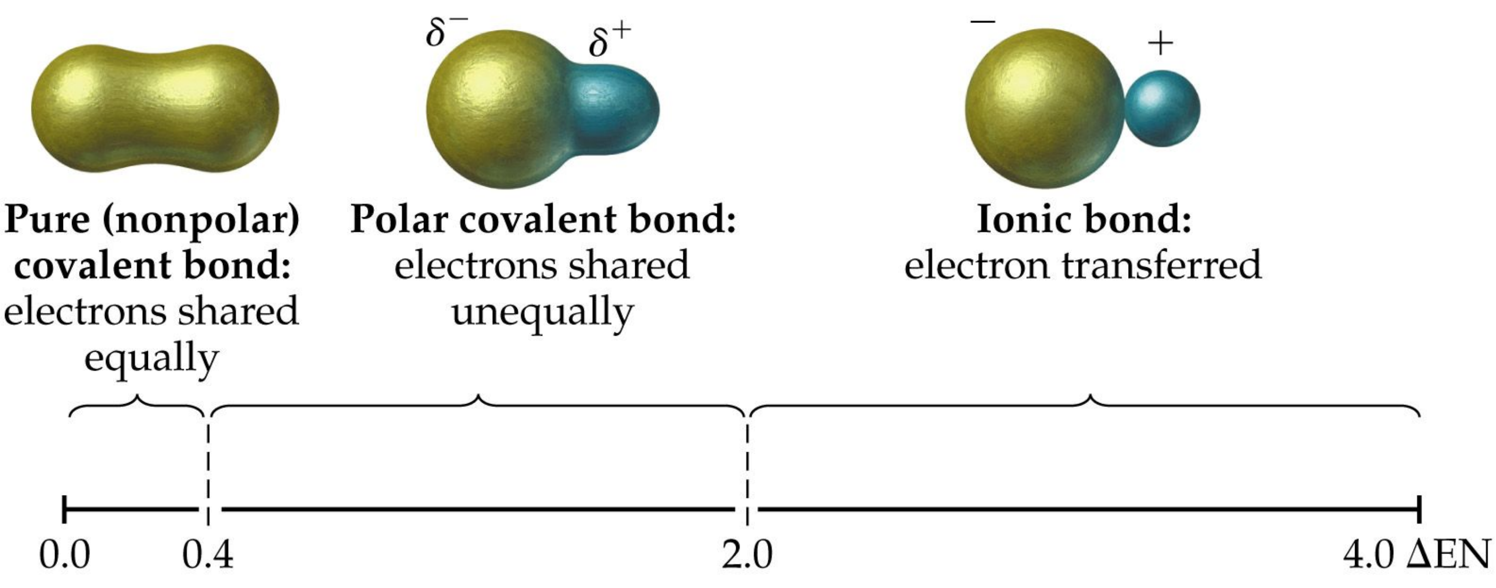
- A measure of the relative tendency of an atom to attract electrons to itself when it is bonded to another atom.
- “Electron Greed”

- Electronegativity increases up and to the right on the periodic table.


IA
H 2.1
IIA
Li 1.0
Be 1.5
Na 0.9
Mg 1.2
K 0.8
Ca 1.0
Rb 0.8
Sr 1.0
Cs 0.7
Ba 0.9
IIIA
IVA
VA
VIA
VIIA
B 2.0
C 2.5
N 3.0
O 3.5
F 4.0
Al 1.5
Si 1.8
P 2.1
S 2.5
Cl 3.0
Ga 1.6
Ge 1.8
As 2.0
Se 2.4
Br 2.8
In 1.7
Sn 1.8
Sb 1.9
Te 2.1
I 2.5
Tl 1.8
Pb 1.9
Bi 1.9
Po 2.0
At 2.1
VIIIA
He
Ne
Ar
Kr
Xe
Rn
0.7 0.8 1.0 1.2 1.4 1.6 1.8 2.0 2.2 2.4 2.6 2.8 3.0 3.2 3.4 3.6 3.8 4.0
Weak
Electronegativity
Strong

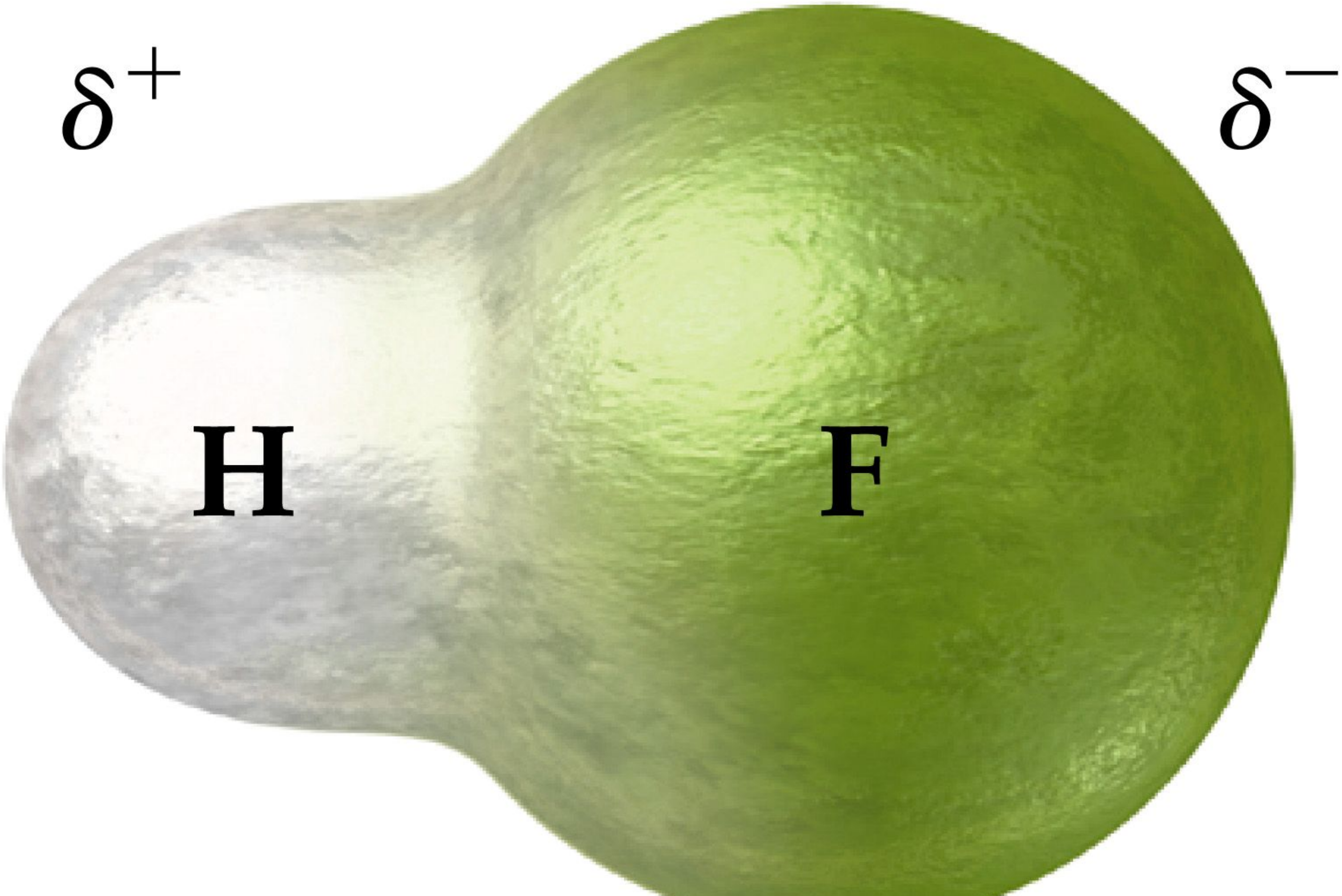
Polar Bonds

- Bonds in which electrons are not shared equally due to the electronegativity differences.



Cl_2

HF



Polar molecules

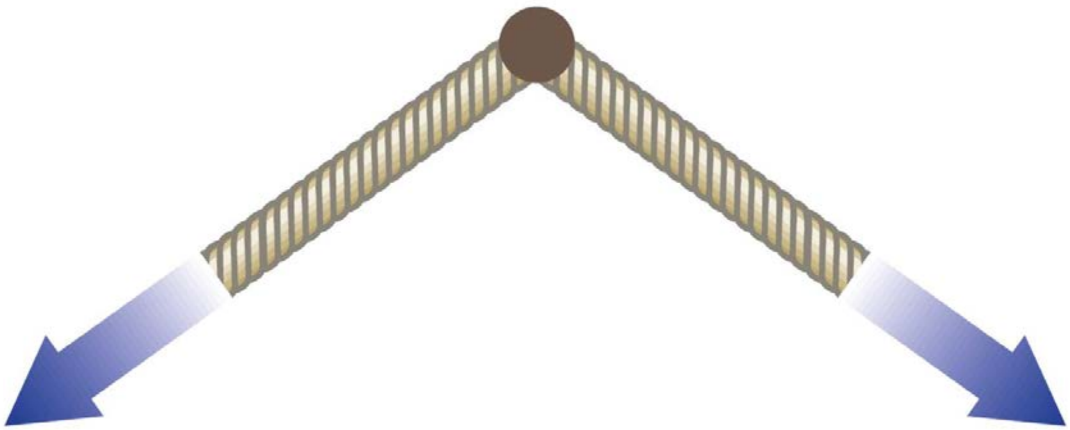
- Molecules with a positive and negative end due to the presence of polar bonds.

Polar Molecules

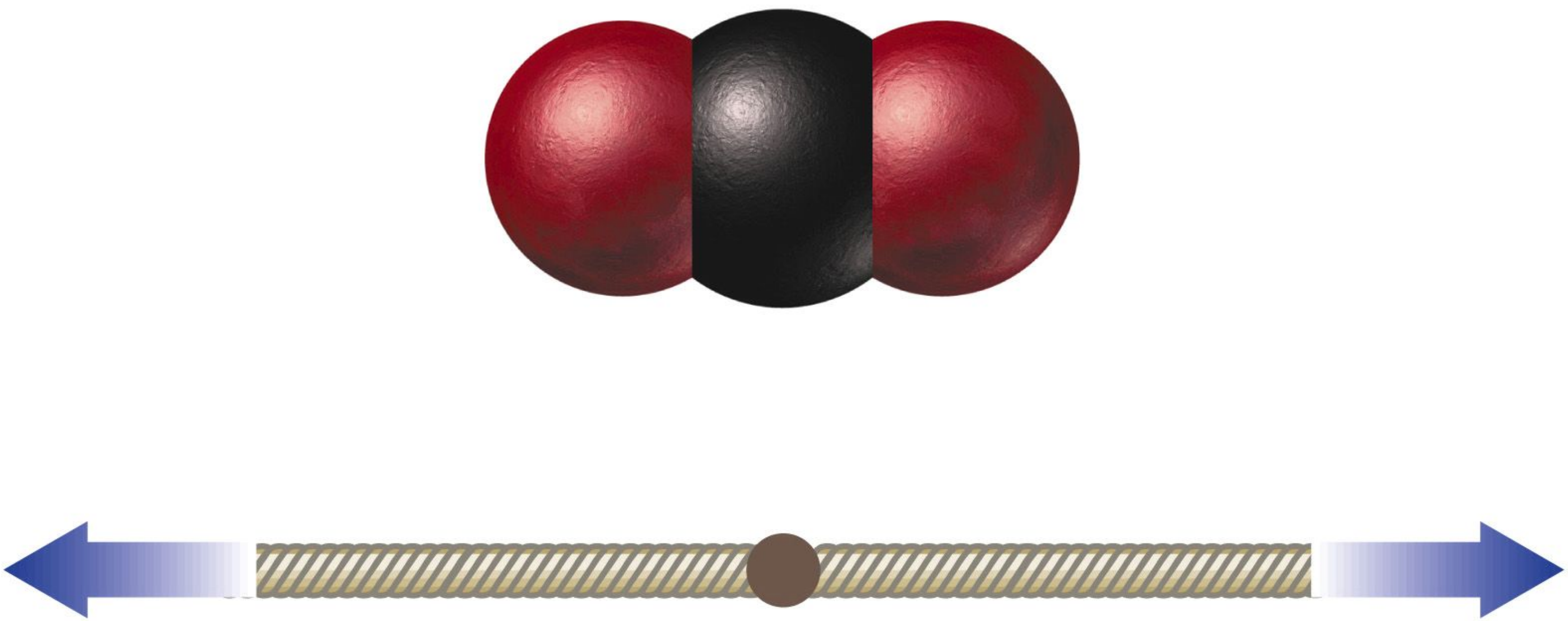
- Dipole - A molecule such as HF which has a positive and a negative end. This dipolar character is often represented by an arrow pointing towards the negative charge.

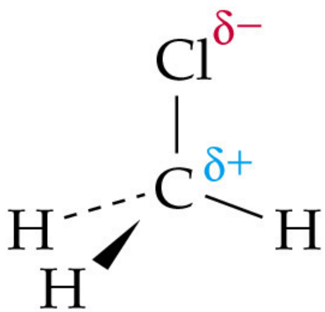
- Dipole moments – the measure of the net molecular polarity
  - Measured in units of Debyes (D) = Qr (charge x separation)

H_2O



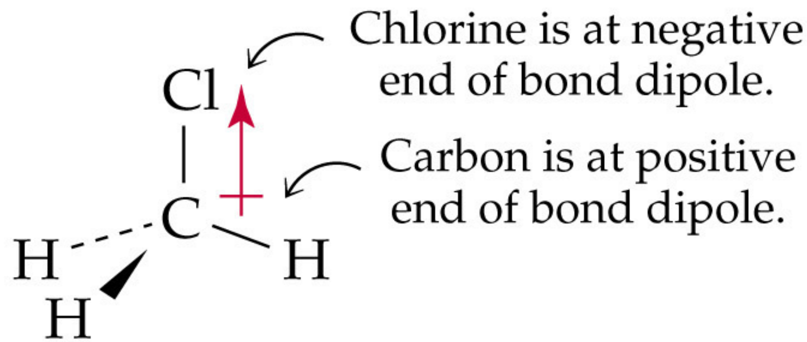
CO_2

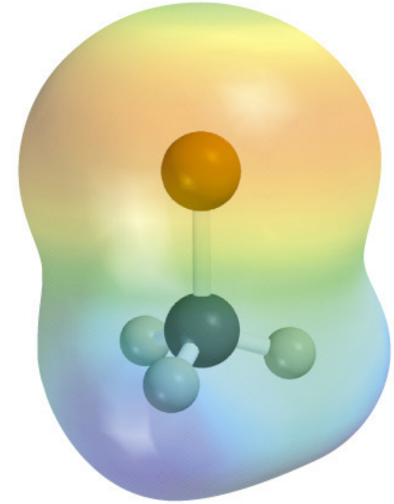


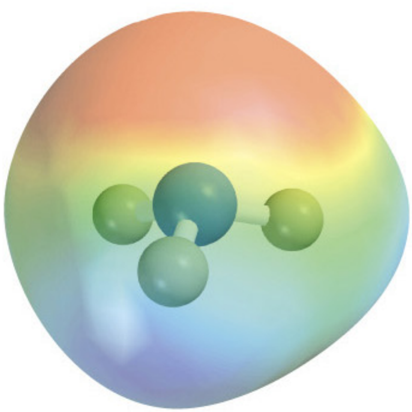


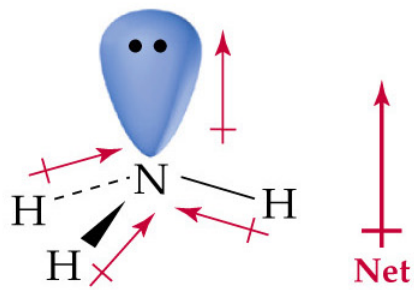
Chloromethane, CH_3Cl

or

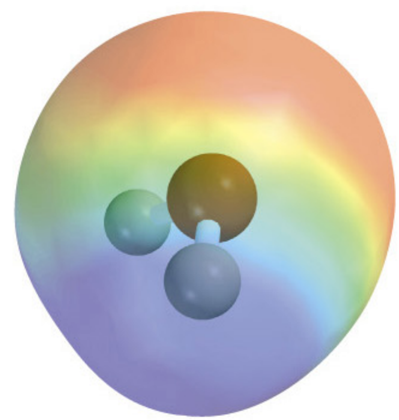


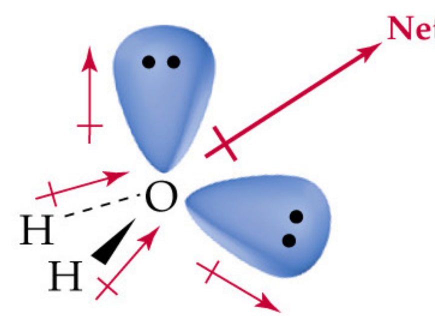






Ammonia (μ = 1.47 D)





Water (μ = 1.85 D)