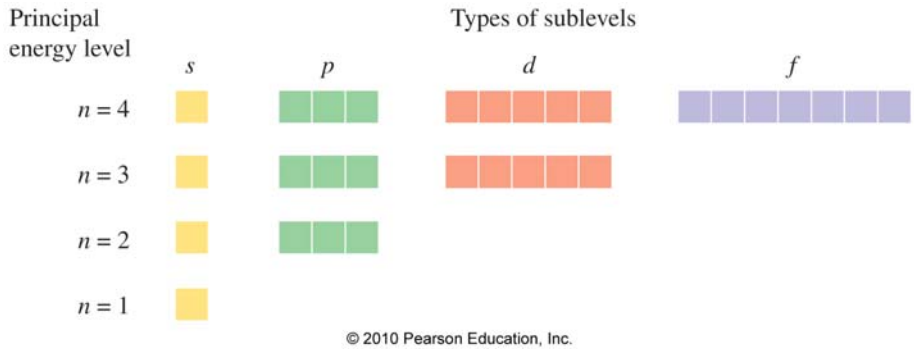
Chapter 5 Compounds and Their Bonds

An **octet refers to 8 valence electrons**

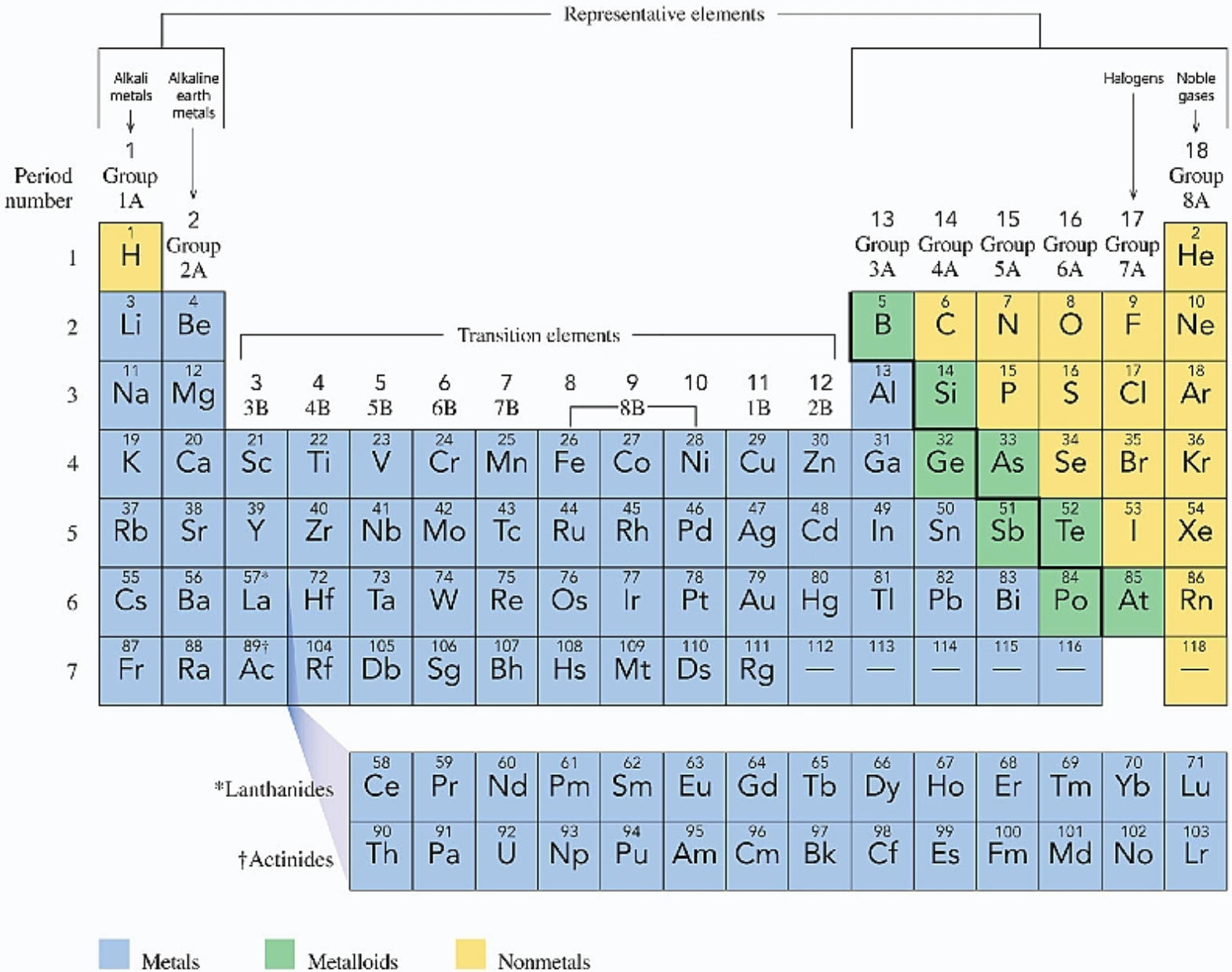
This is associated with the stability of the noble gases other than He;
He is stable with 2 valence electrons (duet)

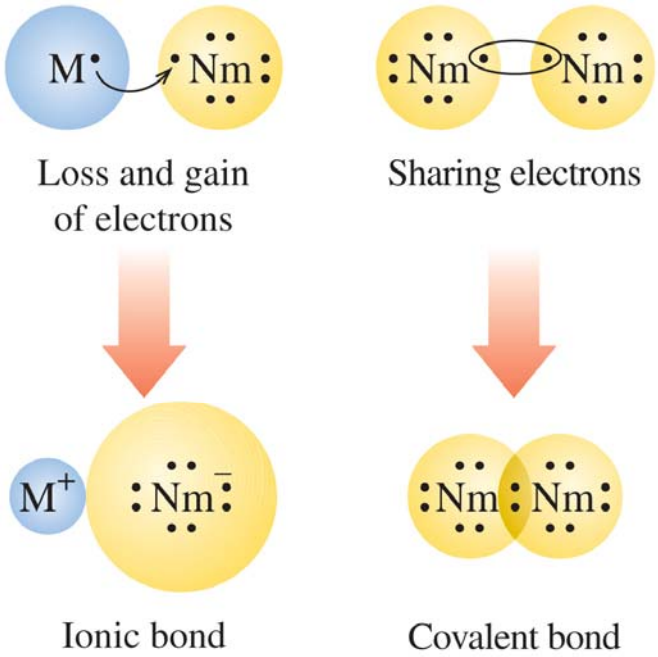
		Valence Electrons
He	$1s^2$	2
Ne	$1s^2\ 2s^2\ 2p^6$	8
Ar	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$	8

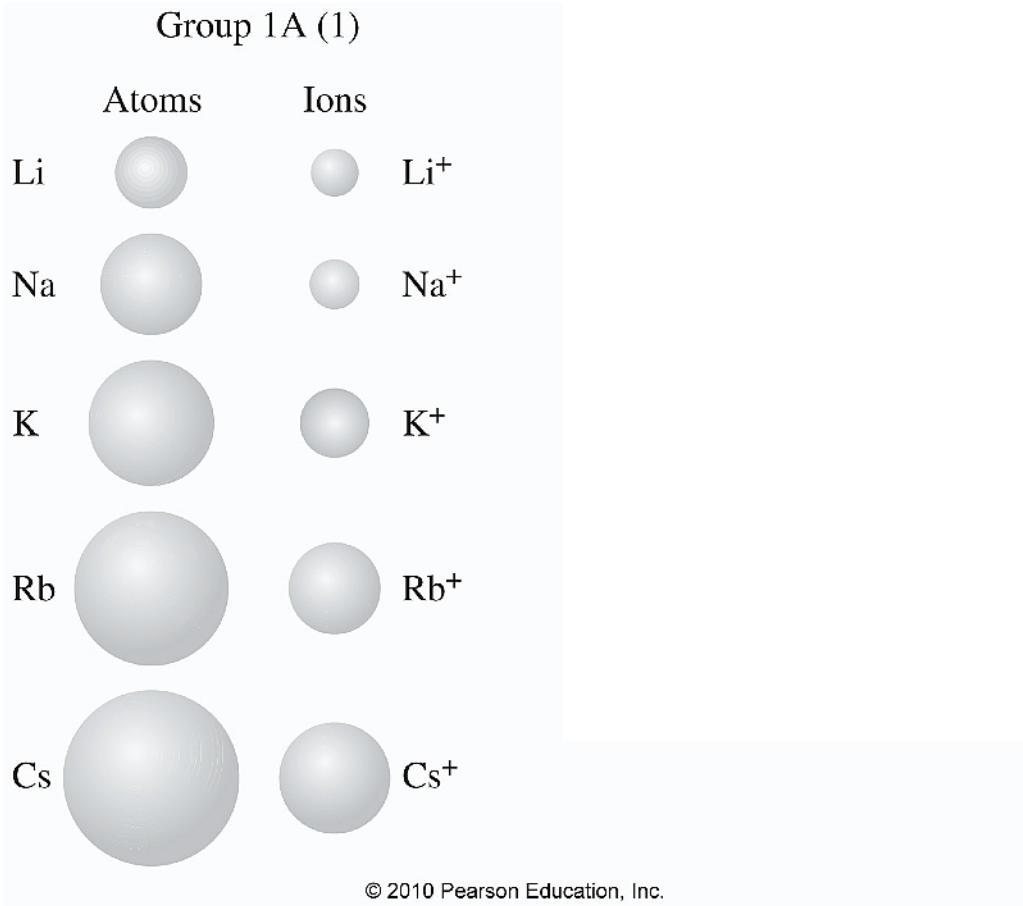


One of the ways chemists have of explaining the formation of both inorganic and organic compounds and their relative stability is to view their reactivity as a attempt to become isoelectronic with either helium or one of the other inert gases

Periodic Table of Elements

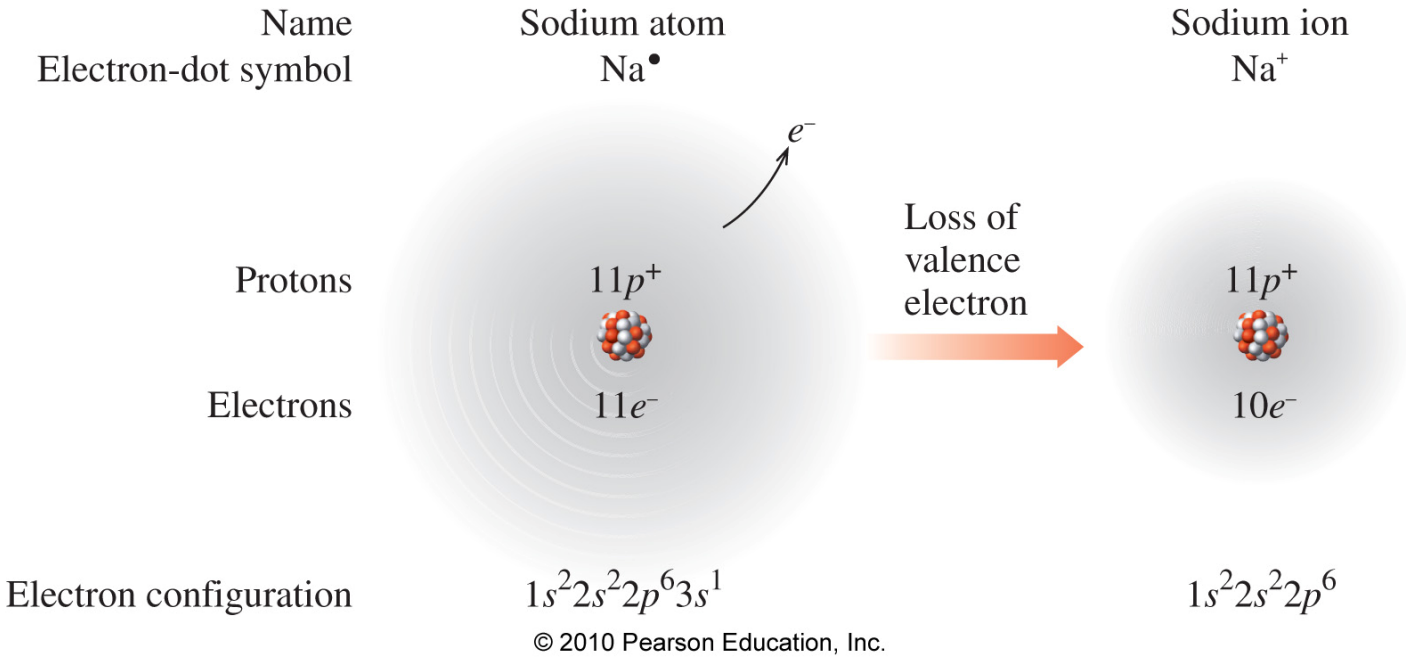


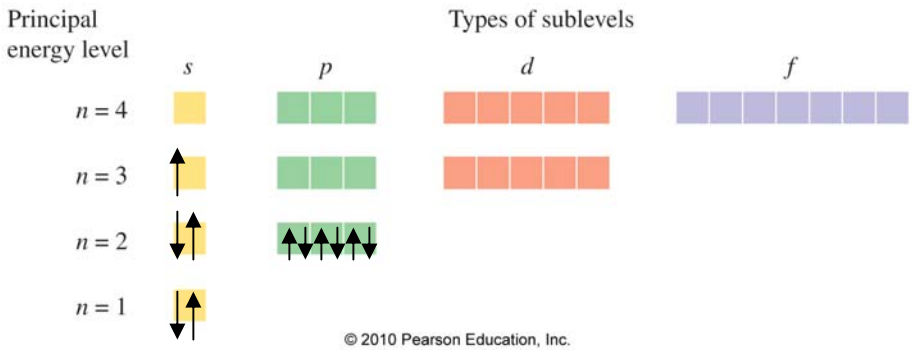


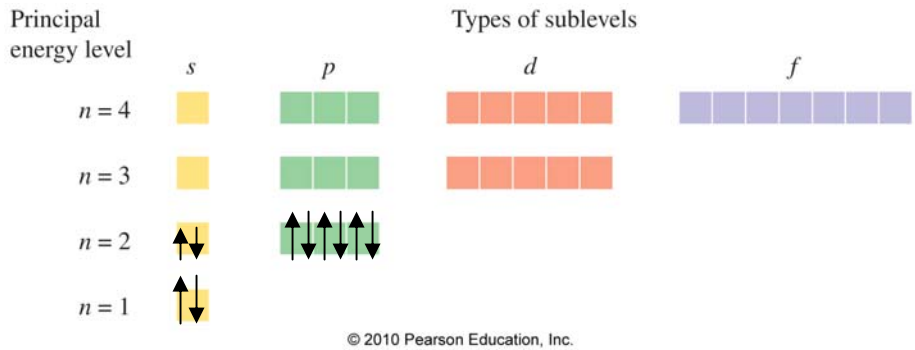


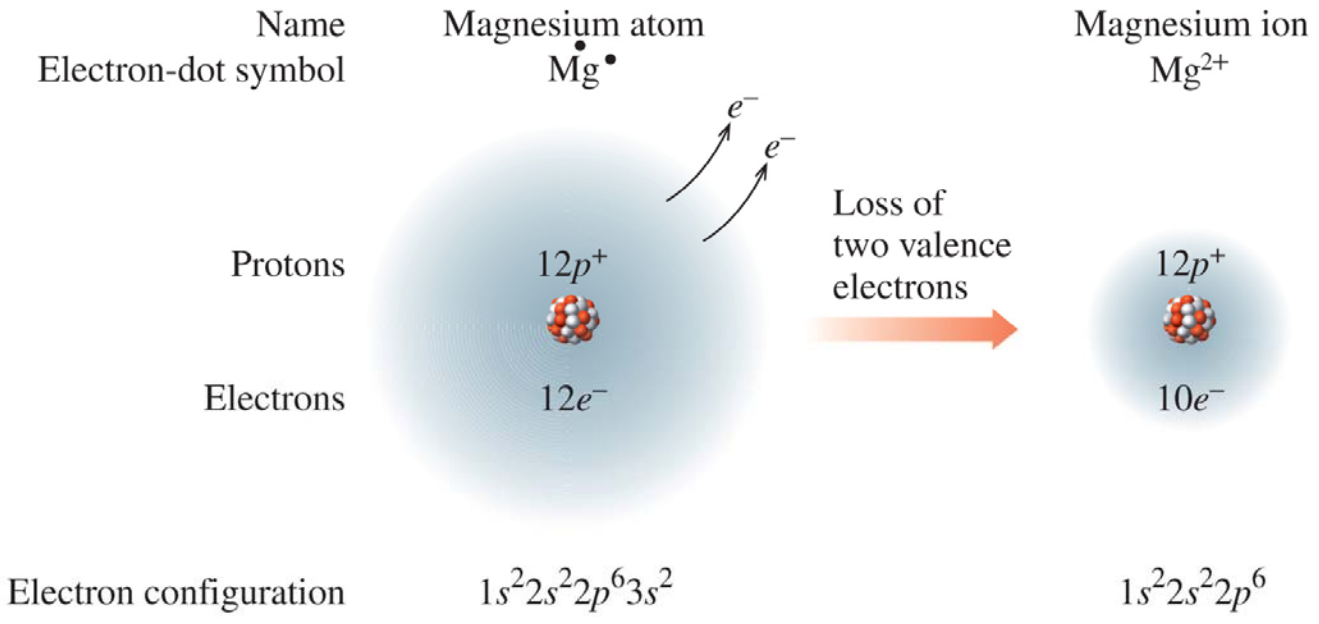
$Li\bullet + :\dot{\underset{\cdot\cdot}{Cl}}: = Li^{+} :\dot{\underset{\cdot\cdot}{Cl}}:^{-1}$ ionic bond

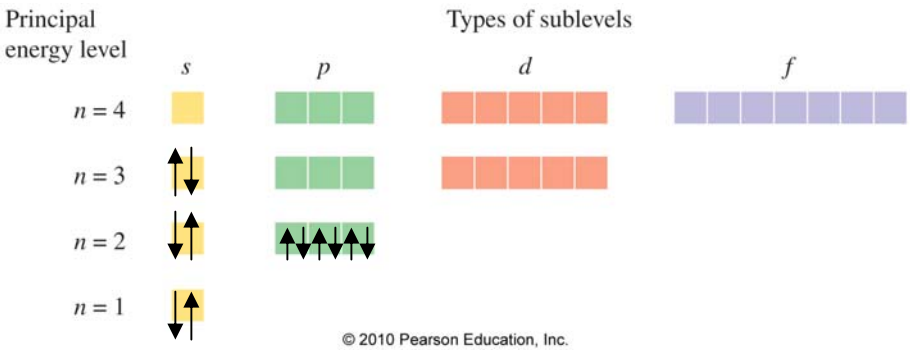
$:\underset{\cdot\cdot}{\overset{\cdot\cdot}{Cl}}\cdot + \cdot\underset{\cdot\cdot}{\overset{\cdot\cdot}{Cl}}: = :\underset{\cdot\cdot}{\overset{\cdot\cdot}{Cl}} : \underset{\cdot\cdot}{\overset{\cdot\cdot}{Cl}}:$ covalent bond











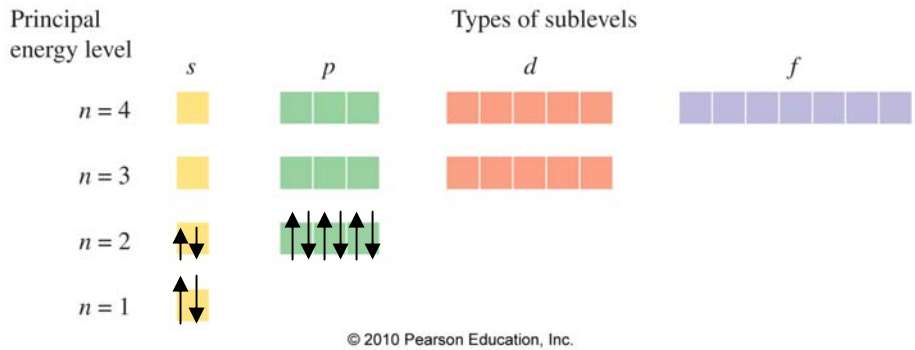
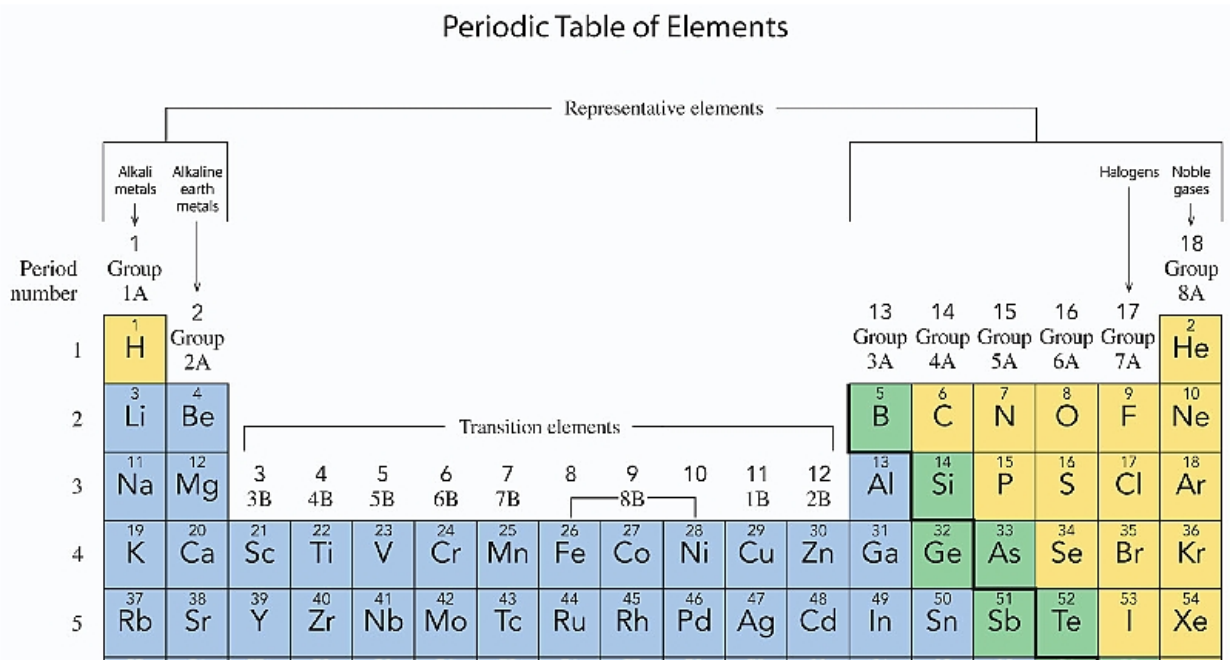
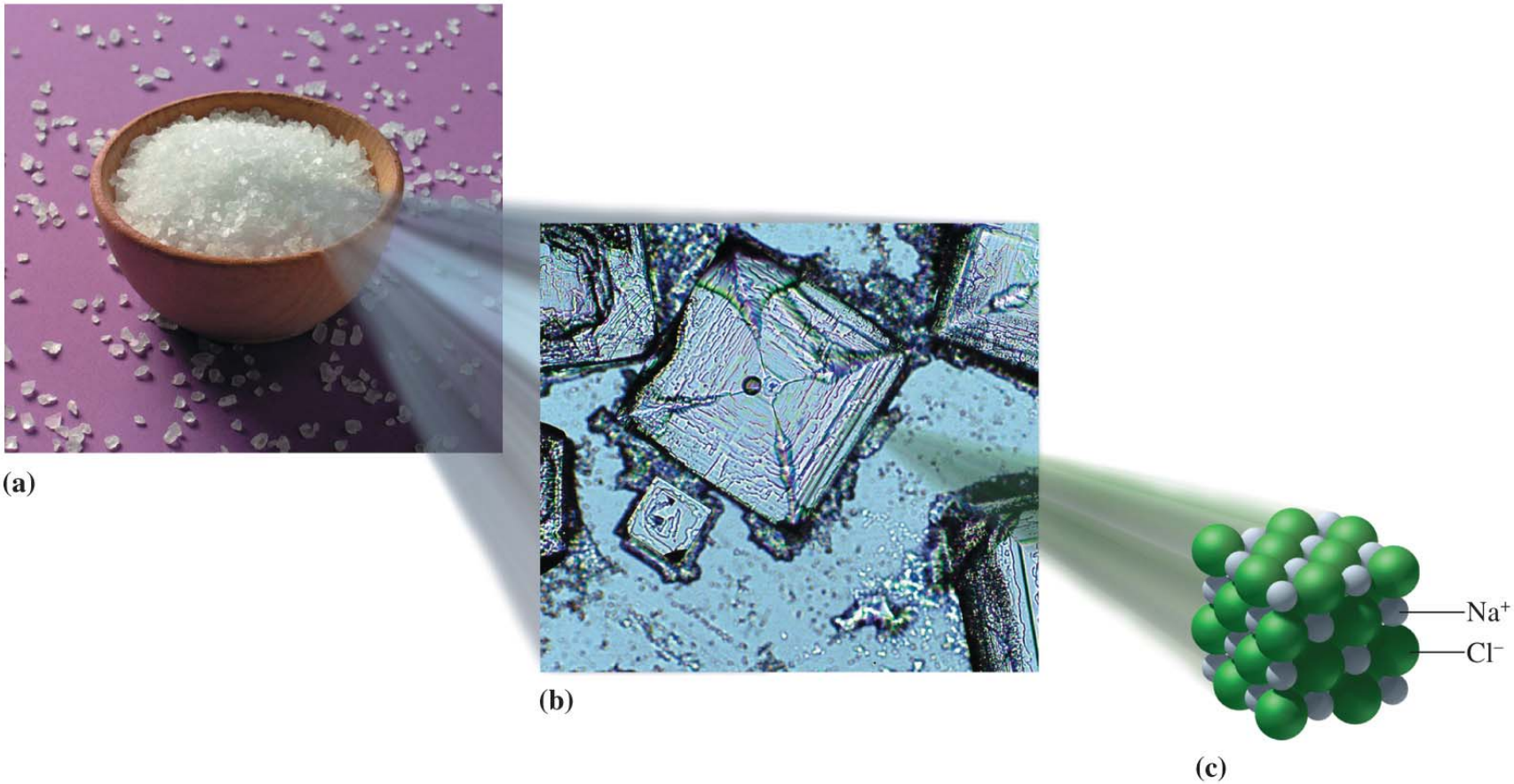


TABLE 5.1 Formulas and Names of Some Common Ions

Group Number	Formula of Ion	Name of Ion	Group Number	Formula of Ion	Name of Ion
	Metals			**Nonmetals**	
1A (1)	Li^+	Lithium	5A (15)	N^{3-}	Nitride
	Na^+	Sodium		P^{3-}	Phosphide
	K^+	Potassium	6A (16)	O^{2-}	Oxide
2A (2)	Mg^{2+}	Magnesium		S^{2-}	Sulfide
	Ca^{2+}	Calcium	7A (17)	F^-	Fluoride
	Ba^{2+}	Barium		Cl^-	Chloride
3A (13)	Al^{3+}	Aluminum		Br^-	Bromide
				I^-	Iodide



Sodium chloride, or “table salt,” is an example of an ionic compound

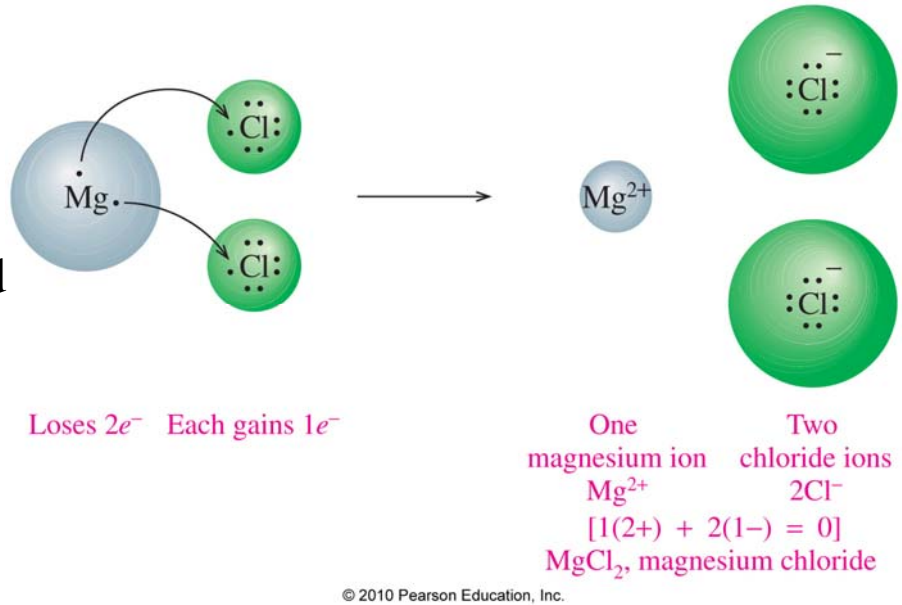


An **ionic formula** consists of positively and negatively charged ions, is neutral, has charge balance

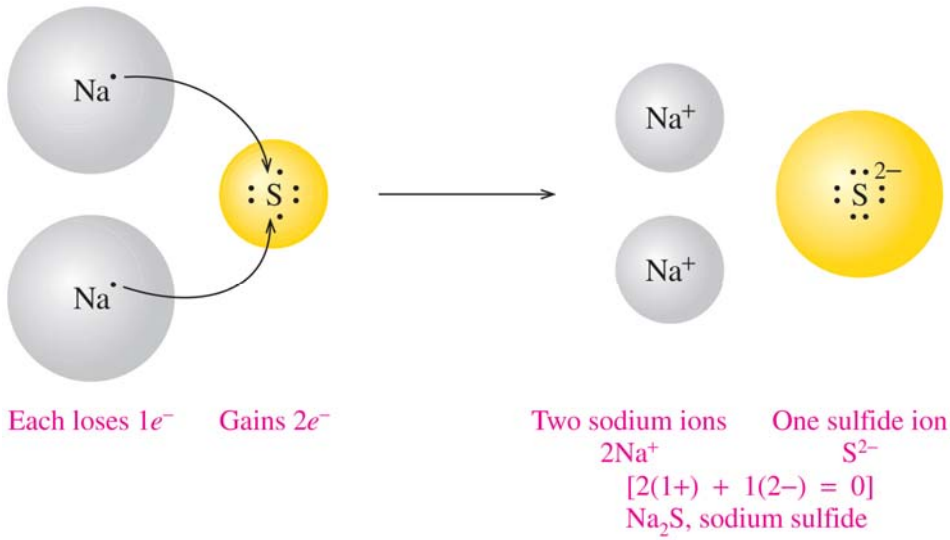
total positive charge = total negative charge

The symbol of the metal is written first, followed by the symbol of the nonmetal.

In **$MgCl_2$**,
a Mg atom loses two valence electrons
two Cl atoms each gain one electron
subscripts indicate the number of ions needed to give charge balance



In **Na_2S**, two Na atoms lose one valence electron;
each S atom gains two electrons;
subscripts show the number of ions needed to give charge balance



Select the correct formula for each of the following ionic compounds:

A. Na^{+} and S^{2-}

1) NaS 2) NaS_2 3 Na_2S

B. Al^{3+} and Cl^{-}

1) $AlCl_3$ 2) AlCl 3) $A1_3Cl$

C. Mg^{2+} and N^{3-}

1) MgN 2) Mg_2N_3 3) Mg_3N_2

Note that Mg^{+2} and Mg^{++} means the same as Mg^{2+}

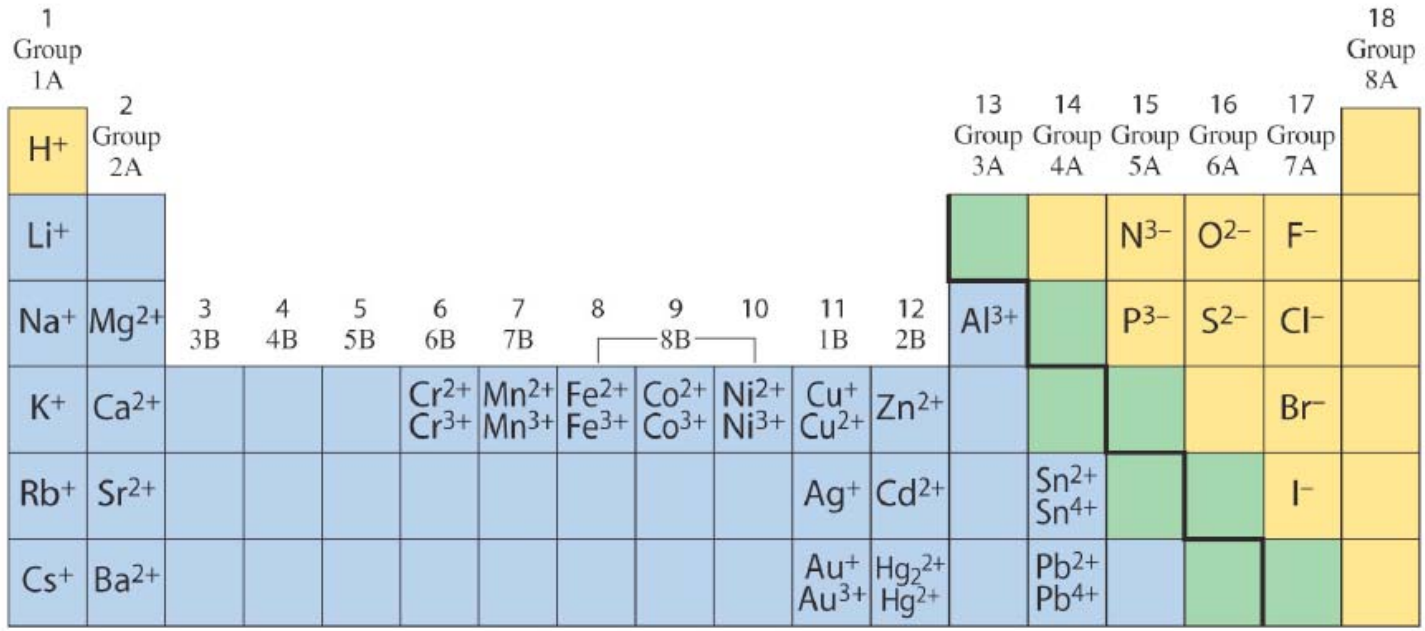
TABLE 5.4 Names of Some Ionic Compounds

Compound	Metal Ion	Nonmetal Ion	Name
NaF	Na^+ Sodium	F^- Fluoride	Sodium fluoride
$MgBr_2$	Mg^{2+} Magnesium	Br^- Bromide	Magnesium bromide
Al_2O_3	Al^{3+} Aluminum	O^{2-} Oxide	Aluminum oxide

Formula	**Ions**		**Name**
	Cation	***Anion***	
NaCl	Na^+	Cl^-	
K_2S	K^+	S^{2-}	
MgO	Mg^{2+}	O^{2-}	
CaI_2	Ca^{2+}	I^-	
Al_2S_3	Al^{3+}	S^{2-}	

Most **transition metals f**orm 2 or more positive ions

Zn^{2+}, Ag^{+}, and Cd^{2+} form only one ion.



The names of transition metals with two or more positive ions (cations) use a *Roman numeral* after the name of the metal to identify ionic charge.

TABLE 5.6 Some Ionic Compounds of Metals That Form Two Kinds of Positive Ions

Compound	Systematic Name
$FeCl_2$	Iron(II) chloride
Fe_2O_3	Iron(III) oxide
Cu_3P	Copper(I) phosphide
$CuBr_2$	Copper(II) bromide
$SnCl_2$	Tin(II) chloride
PbS_2	Lead(IV) sulfide

TABLE 5.5 Some Metals That Form More Than One Positive Ion

Element	Possible Ions	Name of Ion
Chromium	Cr^{2+}	Chromium(II)
	Cr^{3+}	Chromium(III)
Copper	Cu^{+}	Copper(I)
	Cu^{2+}	Copper(II)
Gold	Au^{+}	Gold(I)
	Au^{3+}	Gold(III)
Iron	Fe^{2+}	Iron(II)
	Fe^{3+}	Iron(III)
Lead	Pb^{2+}	Lead(II)
	Pb^{4+}	Lead(IV)
Tin	Sn^{2+}	Tin(II)
	Sn^{4+}	Tin(IV)

Select the correct name for each.

A. Fe_2S_3

1) iron sulfide
2) iron(II) sulfide
3) iron(III) sulfide

B. CuO

1) copper oxide
2) copper(I) oxide
3) copper(II) oxide

The correct formula for each of the following is:

A. copper(I) nitride

1) CuN 2) CuN_3 3) Cu_3N

B. lead(IV) oxide

1) PbO_2 2) PbO 3) Pb_2O_4

Polyatomic Ions

A **polyatomic ion**
is a group of atoms
has an overall ionic charge

Examples:

NH_4^+	ammonium
OH^-	hydroxide
NO_3^-	nitrate
NO_2^-	nitrite
CO_3^{2-}	carbonate
PO_4^{3-}	phosphate
HCO_3^-	hydrogen carbonate (bicarbonate)

Fertilizer
$NaNO_3$

Plaster molding
$CaSO_4$

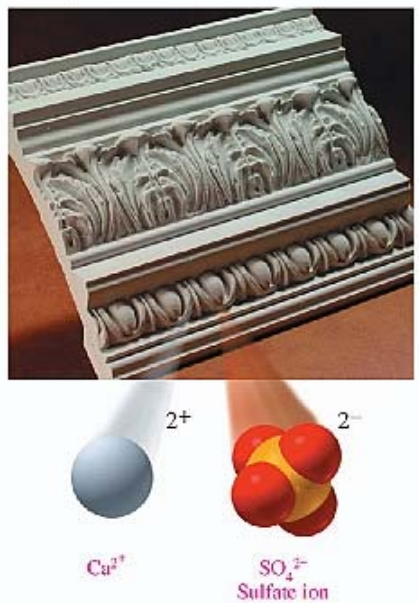


TABLE 5.8 Some Compounds That Contain Polyatomic Ions

Formula	Name	Use
$BaSO_4$	Barium sulfate	Radiopaque medium
$CaCO_3$	Calcium carbonate	Antacid, calcium supplement
$Ca_3(PO_4)_2$	Calcium phosphate	Calcium replenisher
$CaSO_3$	Calcium sulfite	Preservative in cider and fruit juices
$CaSO_4$	Calcium sulfate	Plaster casts
$AgNO_3$	Silver nitrate	Topical anti-infective
$NaHCO_3$	Sodium bicarbonate *or* Sodium hydrogen carbonate	Antacid
$Zn_3(PO_4)_2$	Zinc phosphate	Dental cements
$FePO_4$	Iron(III) phosphate	Food and bread enrichment
K_2CO_3	Potassium carbonate	Alkalizer, diuretic
$Al_2(SO_4)_3$	Aluminum sulfate	Antiperspirant, anti-infective
$AlPO_4$	Aluminum phosphate	Antacid
$MgSO_4$	Magnesium sulfate	Cathartic, Epsom salts

We will discuss the bonding in these ions after we focus on covalent bonding

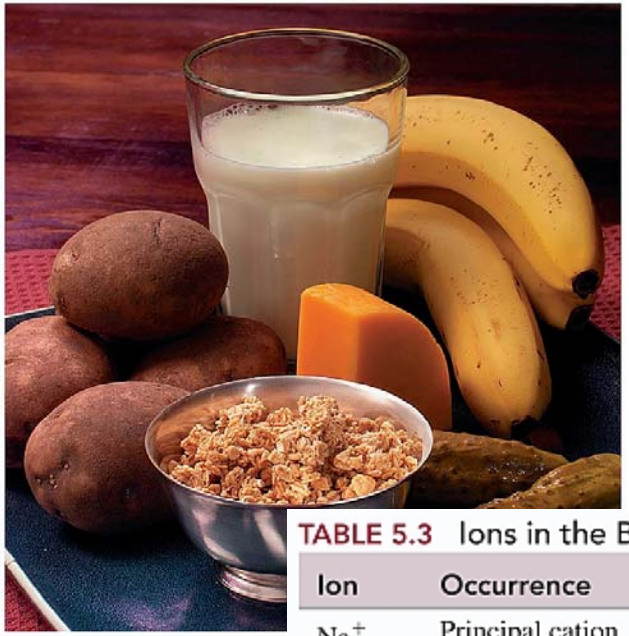
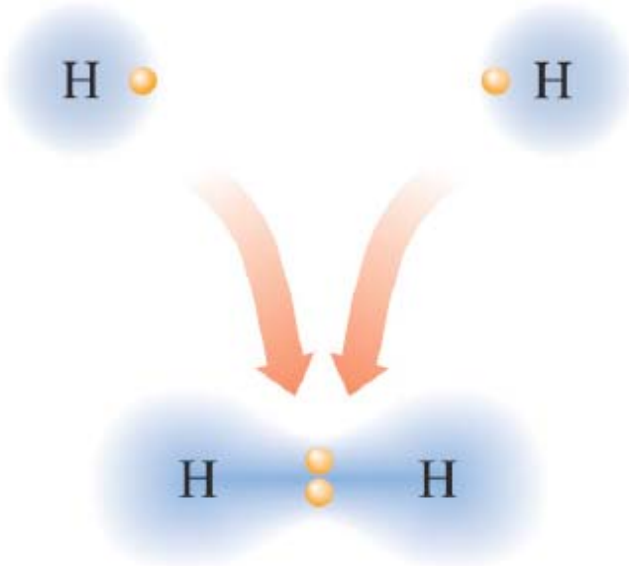


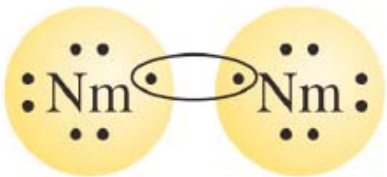
TABLE 5.3 Ions in the Body

Ion	Occurrence	Function	Source	Result of Too Little	Result of Too Much
Na^+	Principal cation outside the cell	Regulation and control of body fluids	Salt, cheese, pickles, potato chips, pretzels	Hyponatremia, anxiety, diarrhea, circulatory failure, decrease in body fluid	Hypernatremia, little urine, thirst, edema
K^+	Principal cation inside the cell	Regulation of body fluids and cellular functions	Bananas, orange juice, milk, prunes, potatoes	Hypokalemia (hypopotassemia), lethargy, muscle weakness, failure of neurological impulses	Hyperkalemia (hyperpotassemia), irritability, nausea, little urine, cardiac arrest
Ca^{2+}	Cation outside the cell; 90% of calcium in the body in bone as $Ca_3(PO_4)_2$ or $CaCO_3$	Major cation of bone; needed for muscle contraction	Milk, yogurt, cheese, greens, spinach	Hypocalcemia, tingling fingertips, muscle cramps, osteoporosis	Hypercalcemia, relaxed muscles, kidney stones, deep bone pain
Mg^{2+}	Cation outside the cell; 70% of magnesium in the body in bone structure	Essential for certain enzymes, muscles, nerve control	Widely distributed (part of chlorophyll of all green plants), nuts, whole grains	Disorientation, hypertension, tremors, slow pulse	Drowsiness
Cl^-	Principal anion outside the cell	Gastric juice, regulation of body fluids	Salt	Same as for Na^+	Same as for Na^+

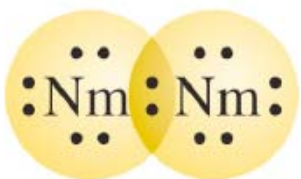
Covalent bonding



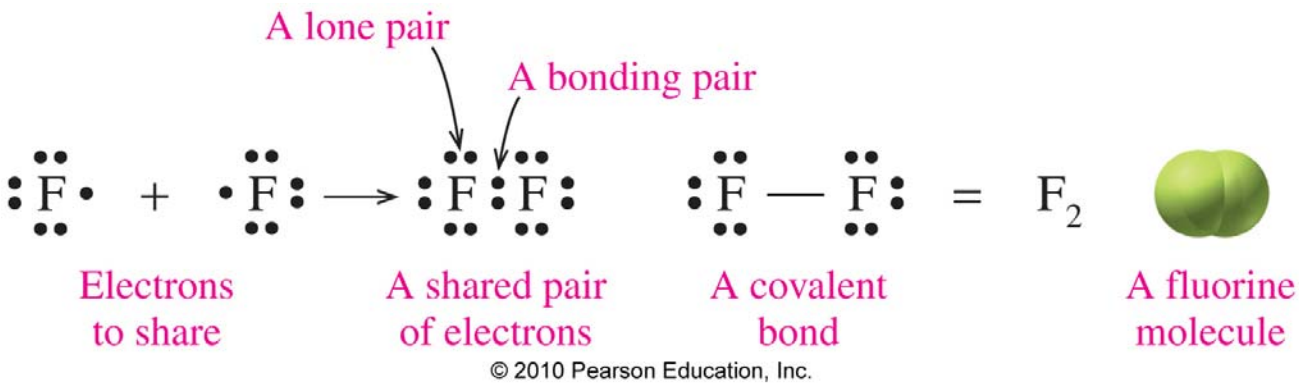
Covalent bonds form
when atoms *share electrons* to complete octets
or duets



Sharing electrons



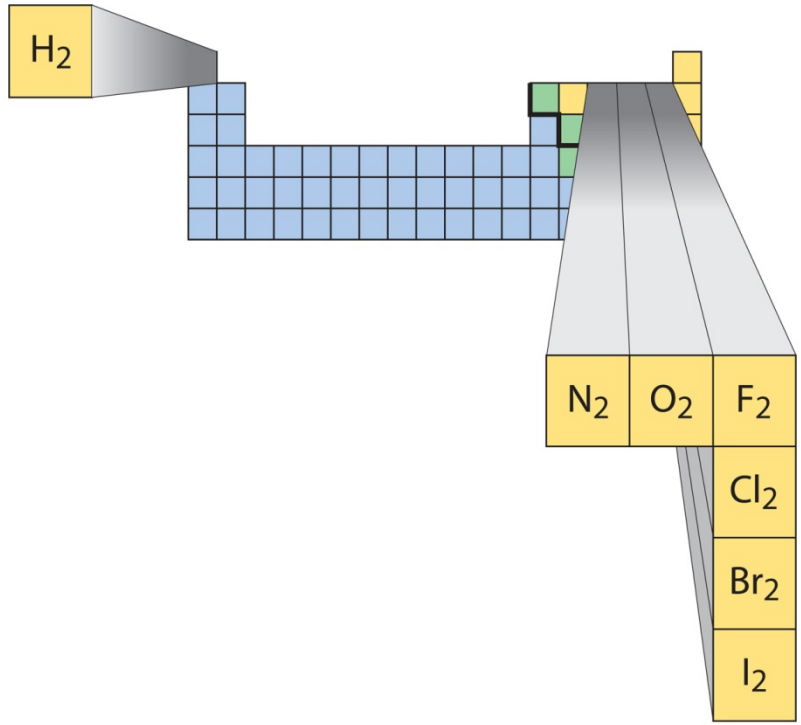
Covalent bond



lone pair electrons refer to 2 electrons not involved in bonding

TABLE 5.9 Elements That Exist as Diatomic, Covalent Molecules

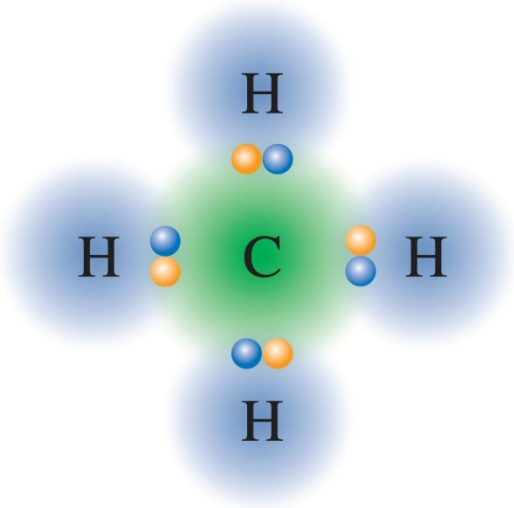
Element	Diatomic Molecule	Name
H	H_2	Hydrogen
N	N_2	Nitrogen
O	O_2	Oxygen
F	F_2	Fluorine
Cl	Cl_2	Chlorine
Br	Br_2	Bromine
I	I_2	Iodine



These are all either gases, liquids or low melting solids compared to the ionic compounds like NaCl and other salts. The interactions between separate molecules is generally weak

Carbon compounds or organic compounds are compounds containing carbon combined with other elements, generally other non-metals

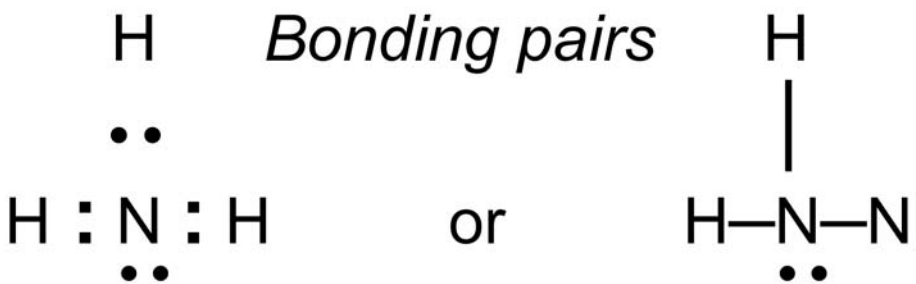
carbon has 4 valence electrons; carbon almost always forms 4 bond, either to itself or to other elements



Methane, CH_4

- Nitrogen has 5 valence electrons. To complete its octet it needs 3 more. It does so generally by sharing three of its own. In NH_3, a N atom is bonded to three H atoms.

The electron-dot structure is written as:



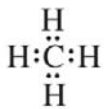
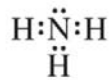
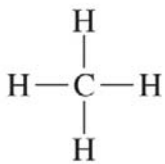
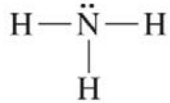
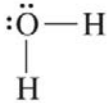
A line (-) in chemistry means the sharing of two electrons

TABLE 5.11 Typical Bonding Patterns of Some Nonmetals in Covalent Compounds

1A (1)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)
[a]H 1 bond					
	[a]B 3 bonds	C 4 bonds	N 3 bonds	O 2 bonds	F 1 bond
		Si 4 bonds	P 3 bonds	S 2 bonds	Cl, Br, I 1 bond

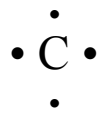
[a]H and B do not form eight-electron octets. H atoms share one electron pair; B atoms share three electron pairs for a set of 6 electrons.

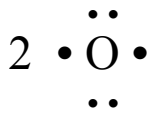
TABLE 5.10 Electron-Dot Formulas for Some Covalent Compounds

CH_4	NH_3	H_2O
Formulas Using Electron Dots Only		
		
Formulas Using Bonds and Electron Dots		
		
Molecular Models		
 Methane molecule	 Ammonia molecule	 Water molecule

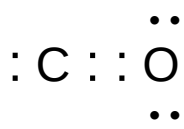
Bonding by the sharing of 4 electrons between atoms

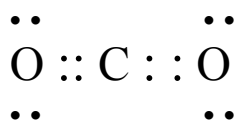
CO_2





How can we arrange these electrons so that every atom has its octet satisfied?





oxygen is satisfied but not carbon; but carbon still has 2 electrons it can use

every atom has an octet:

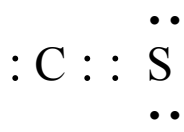
also written as O=C=O

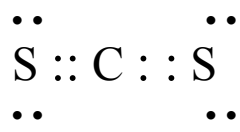
What about sulfur, directly below oxygen in the periodic table?

Bonding by the sharing of 4 electrons between 2 atoms

CS_2

How can we arrange these electrons so that every atom has its octet satisfied?



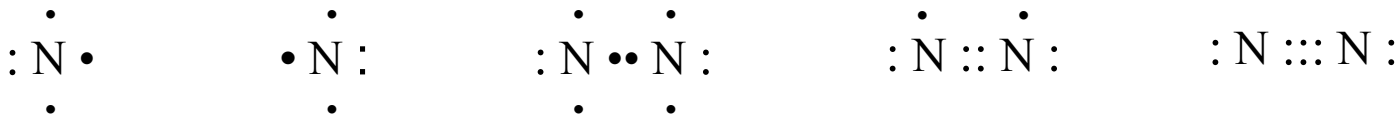


sulfur is satisfied but not carbon; but carbon still has 2 electrons it can use

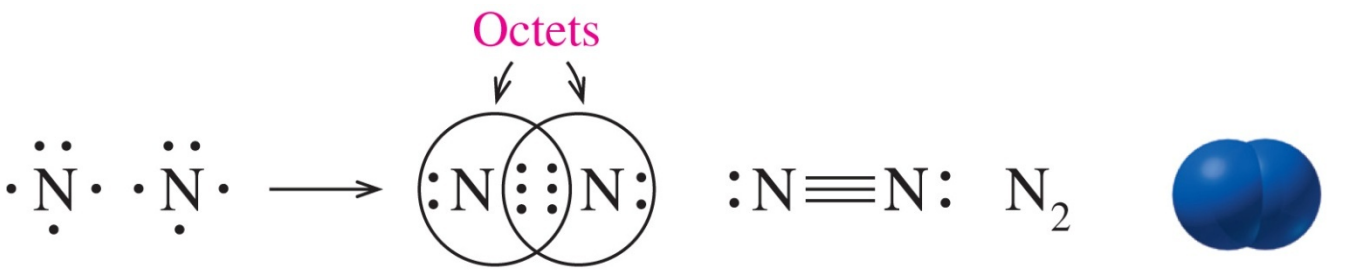
every atom has an octet;

also written as S=C=S

Bonding by the sharing of 6 electrons between 2 atoms, N_2



In a nitrogen molecule, N_2,
each N atom shares 3 electrons
each N atom attains an octet
the sharing of 3 sets of electrons is a multiple bond called a triple bond



Naming and Writing Covalent Formulas

NO	**nitrogen oxide**
NO_2	**nitrogen dioxide**
N_2O_4	**dinitrogen tetroxide**

Guide to Naming Covalent Compounds with Two Nonmetals

STEP 1
Name the first nonmetal by its element name.

STEP 2
Name the second nonmetal by using the first syllable of its element name followed by *ide*.

STEP 3
Add prefixes to indicate the number of atoms (subscripts).

TABLE 5.12 Prefixes Used in Naming Covalent Compounds

1	mono	6	hexa
2	di	7	hepta
3	tri	8	octa
4	tetra	9	nona
5	penta	10	deca

What is the name of SO_3?

STEP 1 The first nonmetal is S sulfur.
STEP 2 The second nonmetal is O, named ox*ide*.
STEP 3 The subscript 3 of O is shown as the prefix ***tri***.

SO_3 → sulfur **tri**oxide

The subscript 1(for S) or *mono* is understood.

TABLE 5.13 Some Common Covalent Compounds

Formula	Name	Commercial Uses
CS_2	Carbon disulfide	Manufacture of rayon
CO_2	Carbon dioxide	Carbonation of beverages; fire extinguishers; propellant in aerosols; dry ice
NO	Nitrogen oxide	Stabilizer
N_2O	Dinitrogen oxide	Inhalation anesthetic: "laughing gas"
SiO_2	Silicon dioxide	Manufacture of glass
SO_2	Sulfur dioxide	Preserving fruits, vegetables; disinfectant in breweries; bleaching textiles
SF_6	Sulfur hexafluoride	Electrical circuits

Select the correct name for each compound.

A. $SiCl_4$

3) silicon tetrachloride

B. P_2O_5

3) diphosphorus pentoxide

C. Cl_2O_7 1) dichlorine heptoxide

Write the correct formula for each of the following:

A. phosphorus pentachloride

1P penta = 5Cl PCl_5

B. dinitrogen trioxide

di = 2N tri = 3 O N_2O_3

C. sulfur hexafluoride

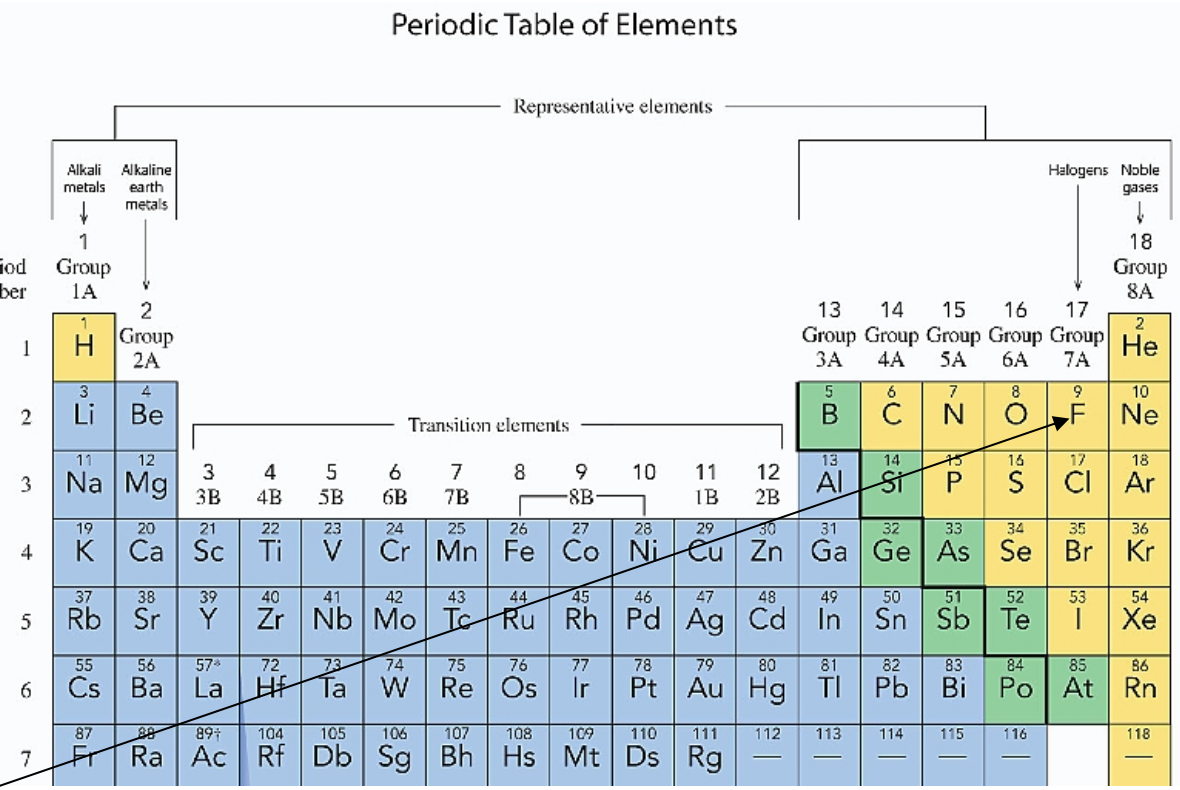
1S hexa = 6F SF_6

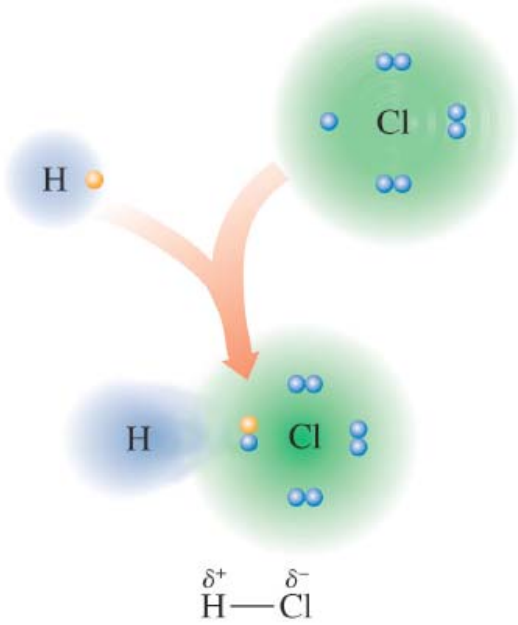
Covalent bonding involves sharing of electrons

This does not guarantee that the electrons are shared equally

An unequal sharing of electron leads to a positive end and a negative end of a molecule called a dipole

Some elements are hogs in regard to sharing





Unequal sharing of electrons in a polar covalent bond (dipole)

The arrow points in the direction of increasing hogishness. Fluorine is the element that is most reluctant to share its electrons. A measure of this is called an element's electronegativity

1 Group 1A	2 Group 2A
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

Electronegativity increases

13 Group 3A	14 Group 4A	15 Group 5A	16 Group 6A	17 Group 7A	18 Group 8A
				H 2.1	
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	

The larger the number, the more hogish the element

TABLE 5.14 Electronegativity Difference and Types of Bonds

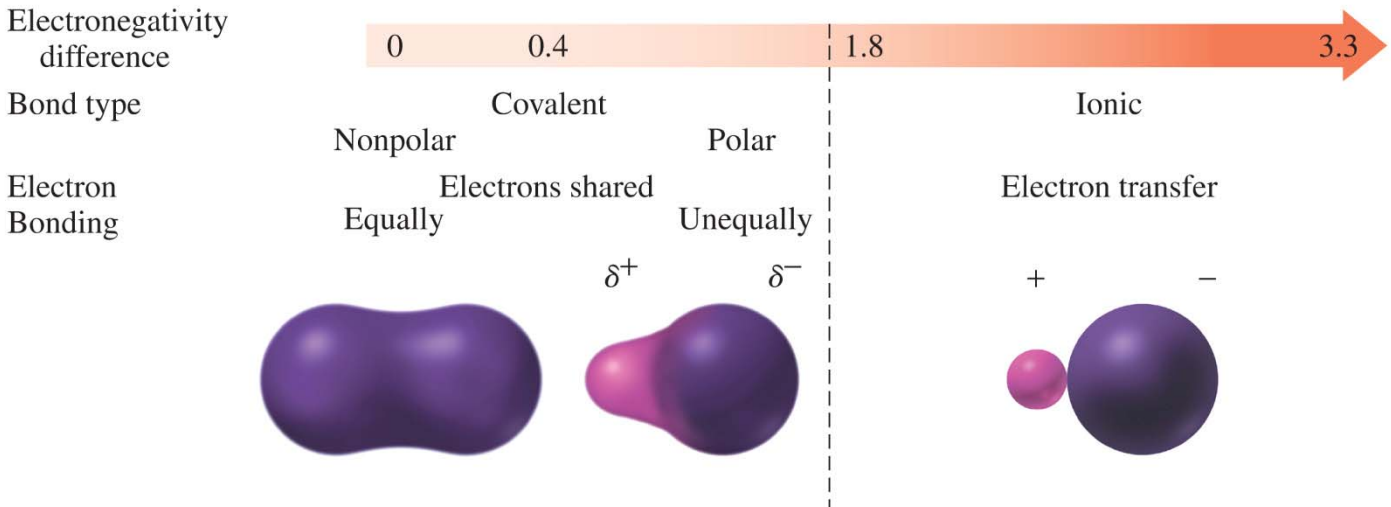


TABLE 5.15 Predicting Bond Type from Electronegativity Differences

Molecule	Bond	Type of Electron Sharing	Bond Type
H_2	H—H	Shared equally	Nonpolar covalent
Cl_2	Cl—Cl	Shared equally	Nonpolar covalent
CH_4	C—H	Shared equally	Nonpolar covalent
HBr	$\overset{\delta^+}{H}—\overset{\delta^-}{Br}$	Shared unequally	Polar covalent
HCl	$\overset{\delta^+}{H}—\overset{\delta^-}{Cl}$	Shared unequally	Polar covalent
NaCl	$Na^+ Cl^-$	Electron transfer	Ionic
MgO	$Mg^{2+} O^{2-}$	Electron transfer	Ionic

[a]Values are taken from Figure 5.5.

Predict if the following are polar or non-polar

F-F

O=C=O

Shapes of Molecules

All diatomic molecules are linear since two points determine a straight line

Suppose we now consider two atoms bonded to a central atom;

Also suppose that the two atom not bonded together, repel each other;

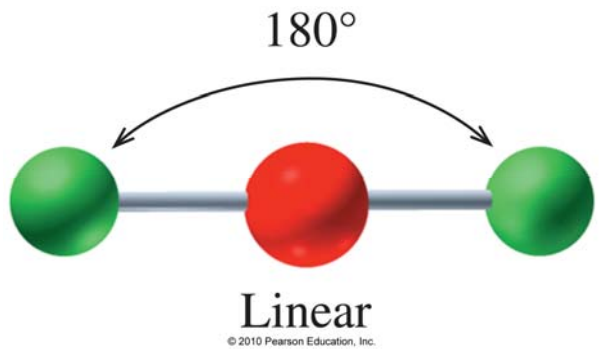
What relative geometry would minimize the repulsion between A and C,

a bent line or a straight line?

A——————B ——————C

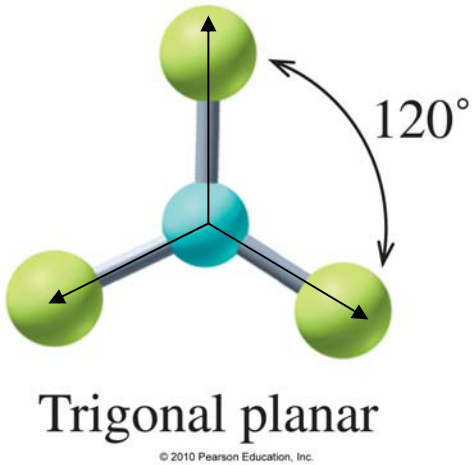
Would it be a polar molecule?

Consider **CO_2**

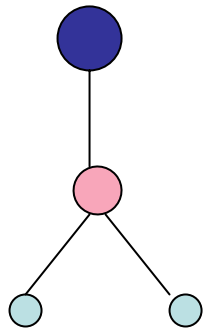


Suppose we now consider three identical atoms bonded to a central atom and that they repel each other;

What relative geometry would minimize the repulsion between the three identical atoms?



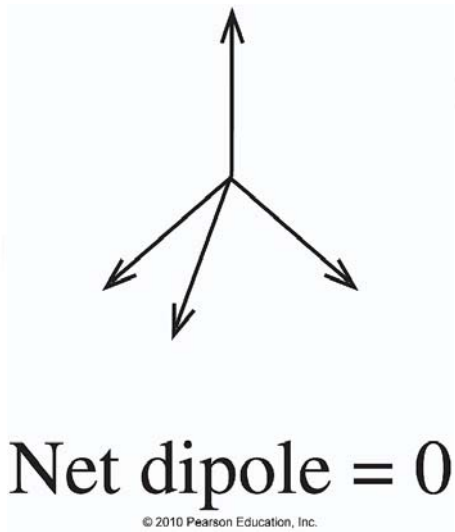
no dipole

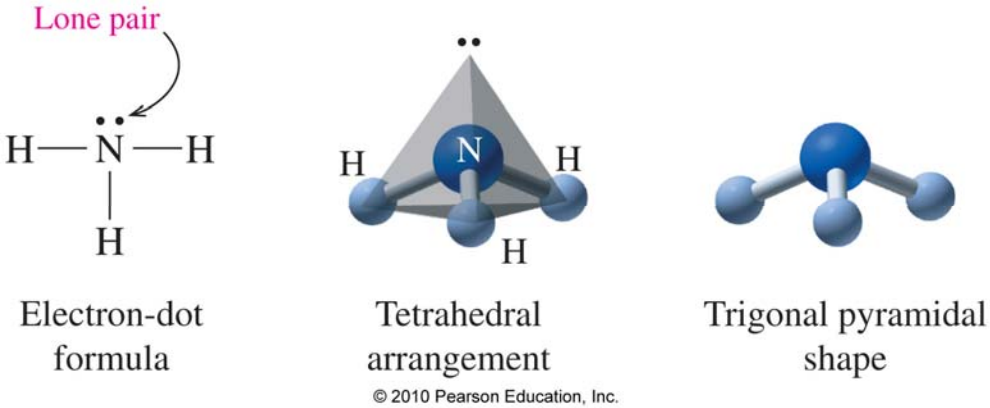


some dipole

What if two of the three atoms are identical and the third one different?

With 4 identical atoms attached to a central atom the following geometry is obtained (eg CH_4)

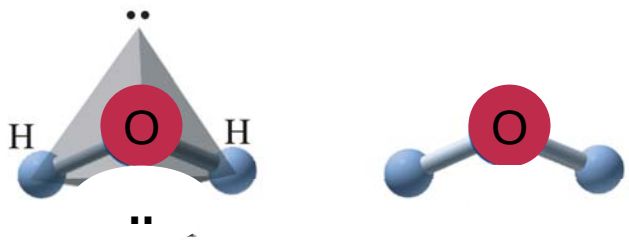




In ammonia, there are three hydrogens and a lone pair of electrons. Do the lone pair of electrons behave as an atom as far as the geometry is concerned?

A lone pair of electrons appears to act as an atom as far as geometry is concerned

What about water which has two lone pairs and two atoms for a total of four groups?



To determine the geometry about a central atom, count the number of groups surrounding the central atom. A group is considered to be either an atom or a lone pair of electrons

2 groups attached results in a linear geometry

3 groups attached results in a triangular geometry

4 groups attached results in a tetrahedral geometry

about the central atom

Shapes of Molecules

There are a number of other geometrical patterns that atoms can form bonded to a central atom. The structures of linear CO_2, pyramidal NH_3, bent H_2O and tetrahedral CH_4 we have just described are the most important to us in this class. Their structure allow us to explain many of their properties.

Consider the following molecules of roughly the same size

water, H_2O ammonia NH_3 methane CH_4

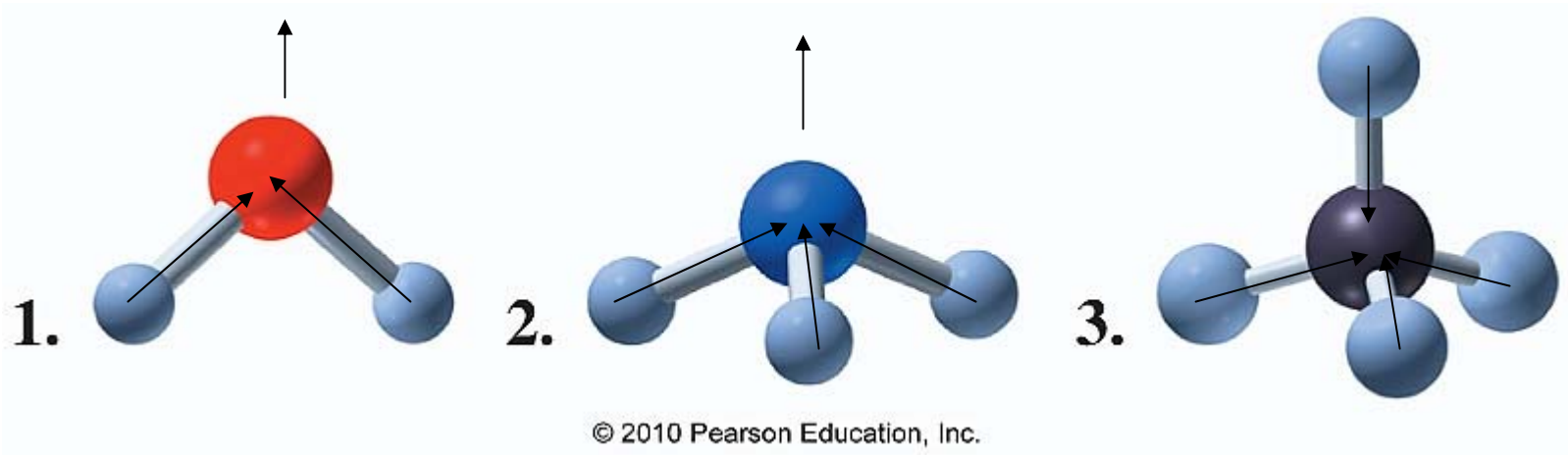
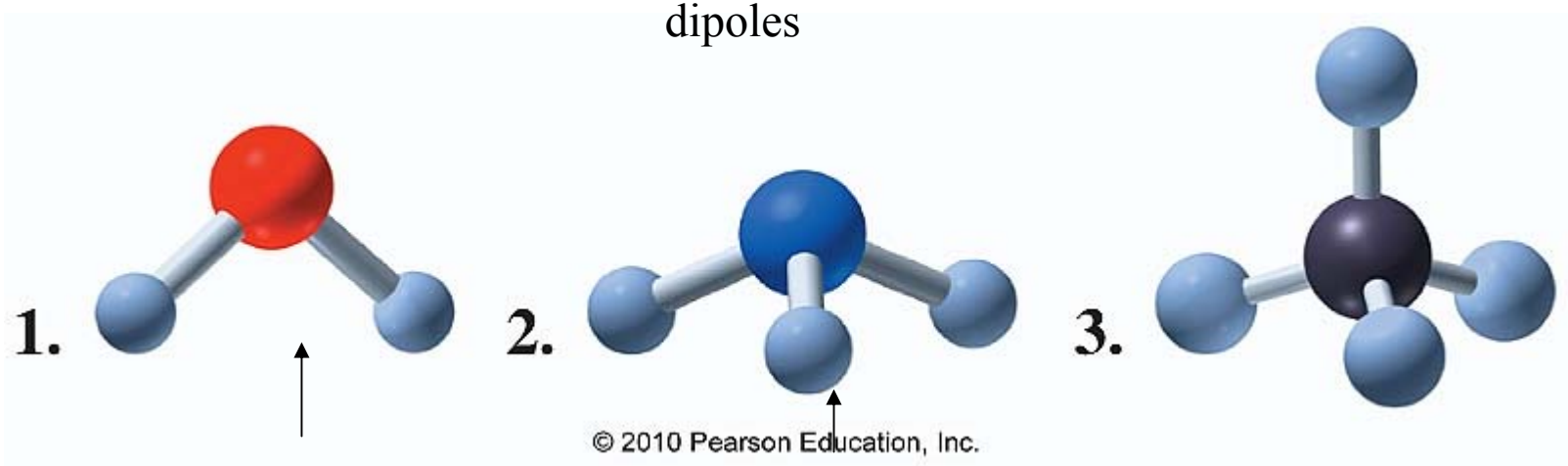


TABLE 2.13 Heats of Fusion and Heats of Vaporization for Selected Substances

Liquid	Formula	Melting Point (°C)	Heat of Fusion (J/g)	Boiling Point (°C)	Heat of Vaporization (J/g)
Water	H_2O	0	334	100	2260
Ethanol	C_2H_5OH	–114	109	78	841
Ammonia	NH_3	–78	351	–33	1380
methane	CH_4			-163	500

dipoles

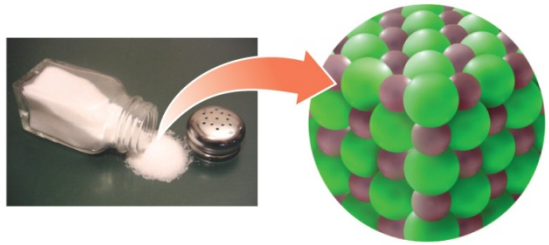


water, H_2O ammonia, NH_3 methane, CH_4

The properties of water are anomalous in comparison to other similar substances; this is due to a property that is referred to as hydrogen bonding which will be discussed later

A summary of the attractive forces between atoms from the strongest to the weakest:

ionic forces in crystals


Sodium chloride

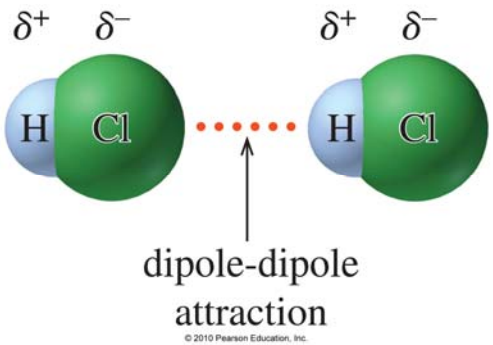
covalent chemical bonds

triple bonds as in nitrogen
double bonds as in CO_2
single bonds as in methane

Hydrogen bonding

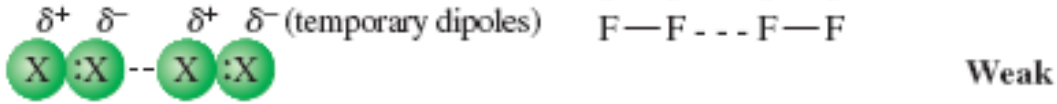
as in H_2O, less in NH_3

dipole-dipole interactions



Dispersion forces are weak attractions between nonpolar molecules caused by temporary (instantaneous) dipoles that develop when electrons are not distributed equally

Dispersion forces
(Temporary shift of electrons in nonpolar bonds)



Weak

The magnitude of these interactions depends on the surface area of the molecule