Professor K

Section 9

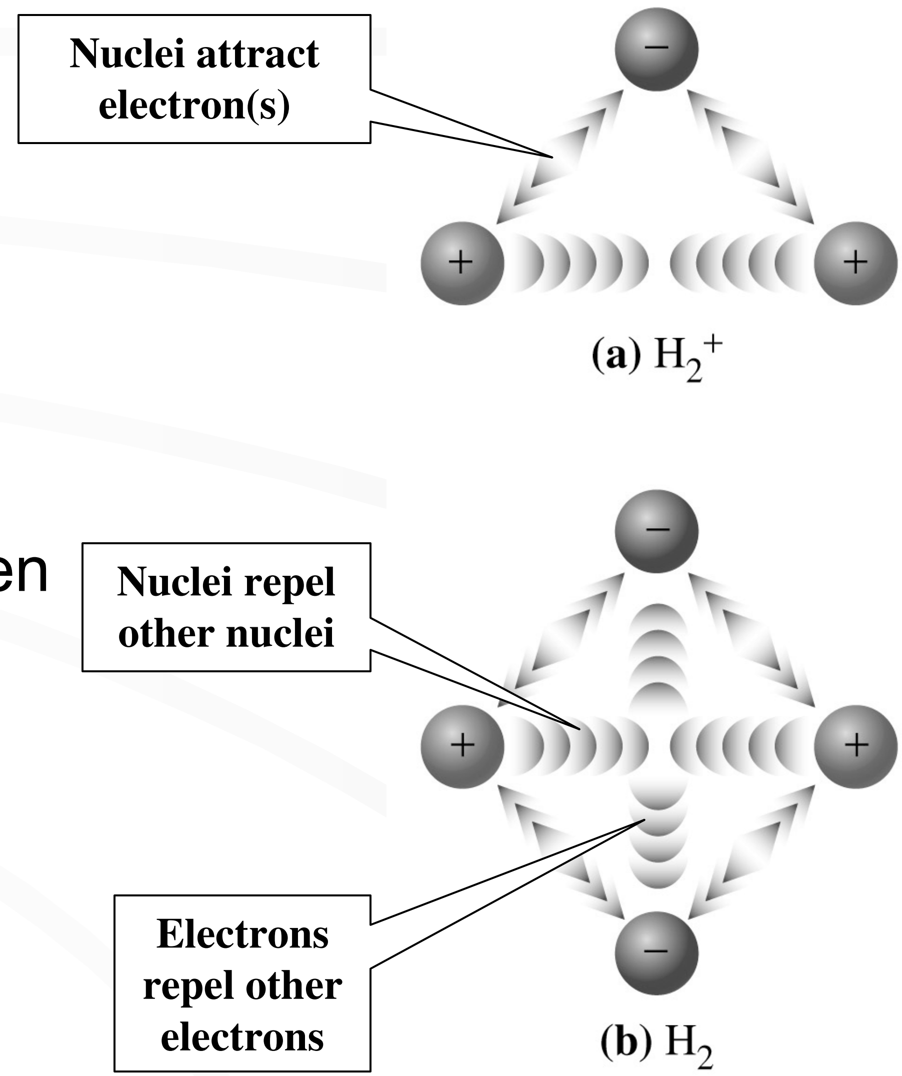
Bonding

Reminder (again)

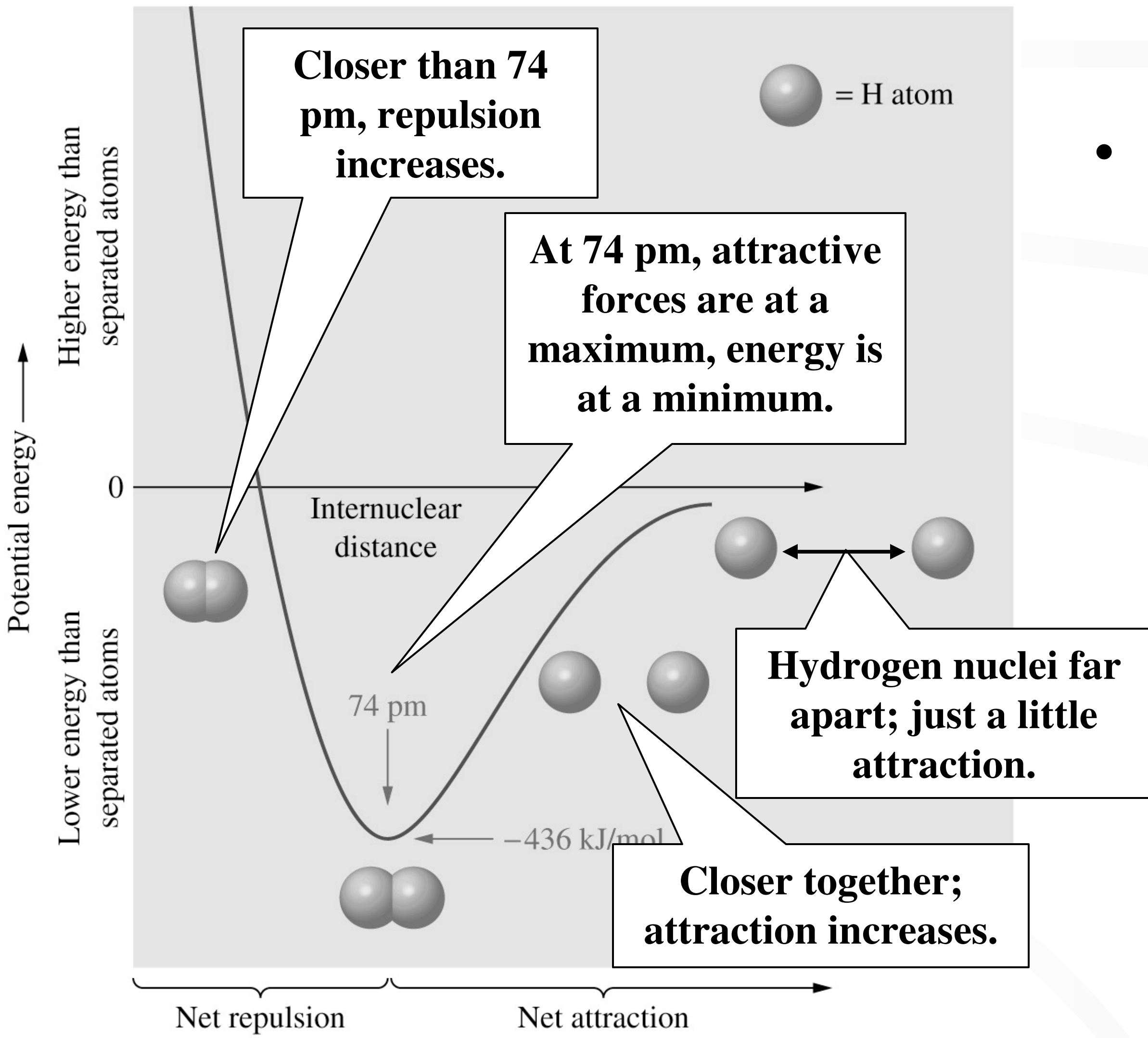
- As you remember, a chemical reaction is the formation and/or breaking of chemical bonds
- A chemical bond is a transfer or sharing of electrons

Chemical bonds

- *Chemical bonds* are the forces that hold atoms together in compounds
  - Bonds are *electrostatic* forces; attractions between opposite charges and repulsions between like charges



Chemical bonds (cont'd)



- *Chemical bonds* are the forces that hold atoms together in compounds
    - When 2 atoms are brought together, there is an *optimum* distance where the unfavorable repulsions and the favorable attractions are maximized (see the *potential energy diagram* at left)

Chemical bonds (cont'd)

- Bonds are responsible for such physical properties as melting point and boiling point (for ionic compounds)

- Bonds (and the type and number of atoms present) ultimately determine the shape of a molecule, and <u>structure determines function</u>

Lewis theory

- Valence electrons participate in bonding
- Metals and nonmetals combine by transferring electrons, forming cations and anions involved in ionic bonds
- Nonmetals combine with each other by sharing electron pairs (overlapping orbitals), forming covalent bonds
- Complete transfer is a 100% ionic "bond"
- Equal sharing is a 100% covalent bond
- The above are two extremes

Lewis theory (cont'd)

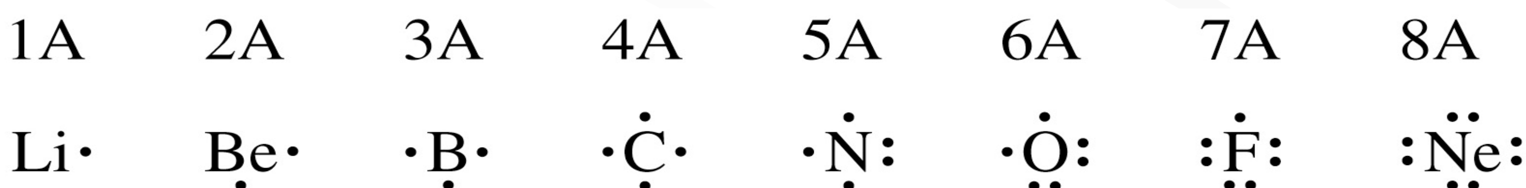
- When atoms gain, lose, or share electrons, atoms tend to acquire the electron configuration of a noble gas ("noble gas configuration")
- Remember, electrons are transferred, NOT protons, so the atom becomes an ion of the original element, not the noble gas itself... noble gas configurations are hyper-stable
- H, Li, and Be follow the duet rule (He configuration), while all other elements (except transition metals) follow the octet rule (Ne, Ar, Kr, Xe, Rn configuration with 8 (or 18) valence electrons)

Lewis symbols

- A Lewis symbol is the element's symbol with "dots" on 4 sides of it to represent the valence electrons
- Electrons are shown "unpaired" whenever possible (although the idea of electron spin had not been developed when Lewis developed his theories)

Lewis symbols (cont'd)

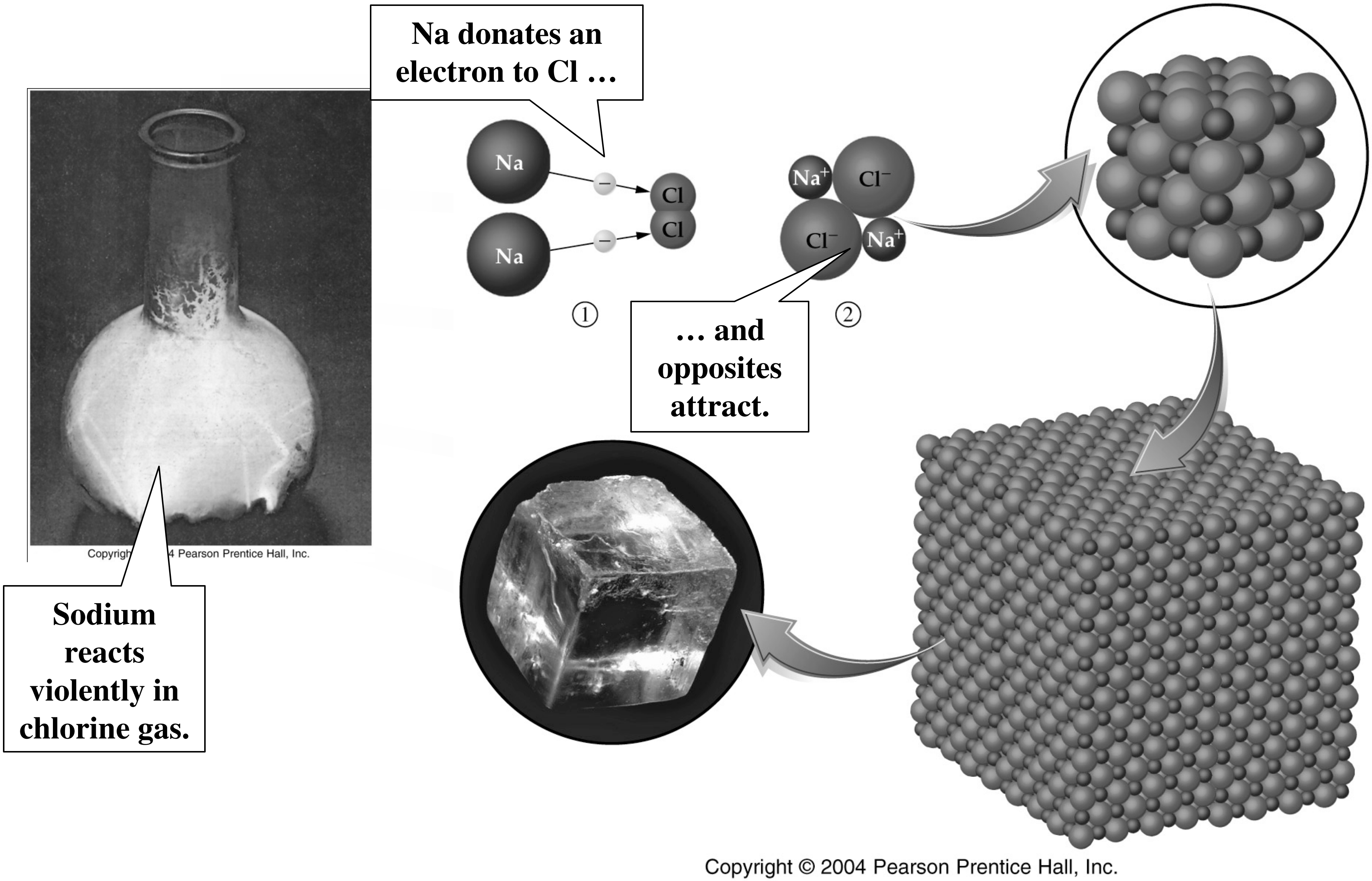
- In a ***Lewis symbol***, the chemical symbol for the element represents the nucleus and core electrons of the atom.
- Dots around the symbol represent the ***valence*** electrons.
- In writing Lewis symbols, the first four dots are placed singly on each of the four sides of the chemical symbol. (Though spin was unknown at the time.)
- Dots are paired as the next four are added.
- Lewis symbols are used primarily for those elements that acquire noble-gas configurations when they form bonds.



Ionic bonds and ionic crystals

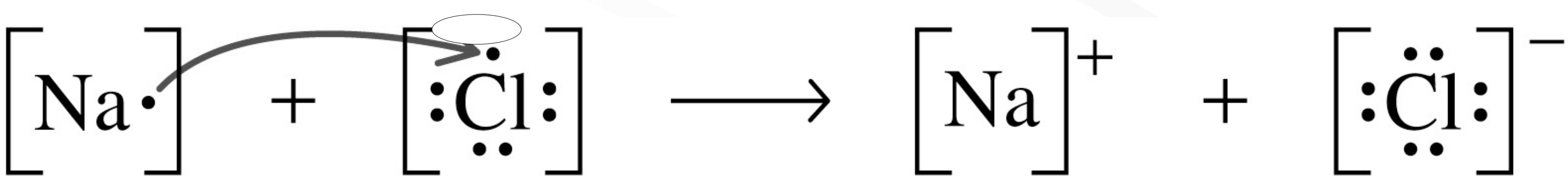
- When atoms lose or gain electrons, they may acquire a noble gas configuration, but do not *become* noble gases.
- Because the two ions formed in a reaction between a metal and a nonmetal have opposite charges, they are strongly attracted to one another and form an ***ion pair***.
- The net attractive electrostatic forces that hold the cations and anions together are ***ionic bonds***.
- The highly ordered solid collection of ions is called an ***ionic crystal***.

Formation of a crystal of sodium chloride



Using Lewis symbols to represent ionic bonding

- Lewis symbols can be used to represent ionic bonding between nonmetals and:
  the s-block metals, some p-block metals, and a few d-block metals.
- Instead of using complete electron configurations to represent the loss and gain of electrons, Lewis symbols can be used.



Example

- Use Lewis symbols to show the formation of ionic bonds between magnesium and nitrogen.
- What are the name and formula of the compound that results?
- (worked out next slide)

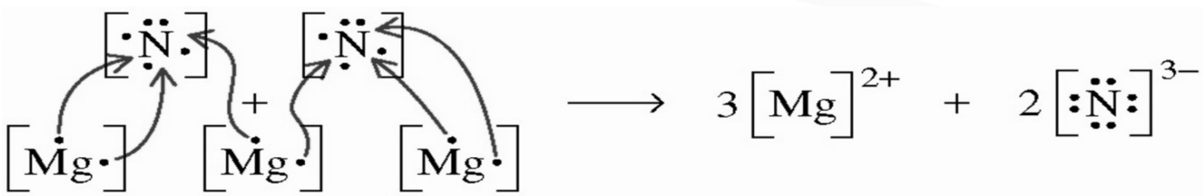
Example worked out

Strategy

First, we determine the number of electrons that must be lost by magnesium atoms and gained by nitrogen atoms so that the resulting ions have noble-gas electron configurations.
Then we can combine Lewis symbols for the ions in the appropriate proportions to produce a neutral formal unit, from which we can deduce the compound's name and formula.

Solution

Mg atoms (group 2A) lose their two valence electrons, and N atoms (group 5A) gain three additional valence electrons. To produce an electrically neutral formula unit, three Mg atoms must lose a total of six electrons and two N atoms must gain a total of six:



The compound is magnesium nitride, Mg_3N_2.

- **Exercise**

  Use Lewis symbols to show the formation of ionic bonds between barium and iodine.
  What are the name and formula of the compound that results?

- **Exercise**

  Use Lewis symbols to show the formation of ionic bonds in aluminum oxide.

Reactions

- Reactions can be drawn showing the transfer of electrons
  - Looking at tables of ionization energy and electron affinity *along with the energies of some changes in state, lattice energies (2 slides ahead) and bond energies*, the energy of a reaction can be calculated (Ex.- the *enthalpy of formation* of NaCl seven slides ahead) just like with Hess' law

Energy changes in ionic compound formation

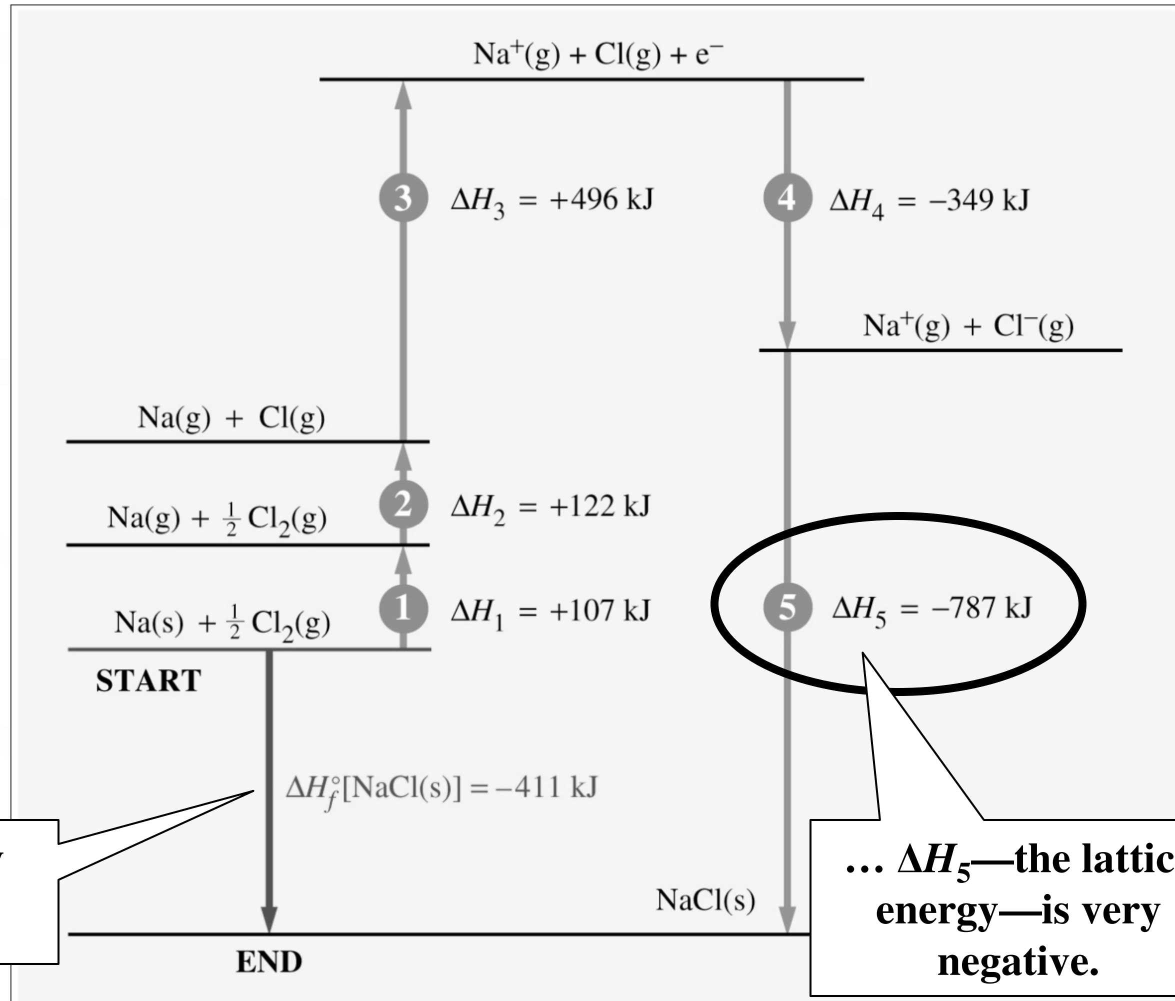
$Na(g) \rightarrow Na^+(g) + e^-$ $\quad IE_1 = +496$ kJ/mol

$Cl(g) + e^- \rightarrow Cl^-(g)$ $\quad EA_1 = -349$ kJ/mol

- From the data above, it doesn't appear that the formation of NaCl from its elements is energetically favored. However …
- … the ***enthalpy of formation*** of the ionic compound is more important than either the first ionization energy or electron affinity.
- The overall enthalpy change can be calculated using a step-wise procedure called the ***Born–Haber cycle***.

Energy changes in ionic compound formation *(cont'd)*

- The Born–Haber cycle is a hypothetical process, in which ΔH_f is represented by several steps.
- What law can be used to find an enthalpy change that occurs in steps??



ΔH_f° for NaCl is very negative because …

… ΔH_5—the lattice energy—is very negative.

Example

- Use the following data to determine the lattice energy of $MgF_2(s)$:
  - enthalpy of sublimation of Mg: +146kJ/mol;
  - I_1 for Mg: +738kJ/mol;
  - I_2 for Mg: +1451kJ/mol;
  - bond-dissociation energy of $F_2(g)$: +159kJ/mol F_2;
  - electron affinity of F: -328 kJ/mol F;
  - enthalpy of formation of $MgF_2(s)$: -1124kJ/mol
- (worked out next slide)

Example worked out

Strategy

The approach we need here differs in three ways from the one used for NaCl(s). (1) The compound MgF_2(s) has two anions for each cation in the crystal. Therefore, in the step where F_2(g) dissociates, we need the bond-dissociation energy based on 1 mol of F_2(g) in order to get 2 mol of F(g). Similarly, in the electron affinity step, we must produce 2 mol of F^-(g) rather than 1 mol. (2) Because the magnesium cation carries a 2+ charge, we must include two ionization steps and both I_1 and I_2 in our calculation. (3) The enthalpy of formation of MgF_2(s) is given, and our unknown is the lattice energy.

Solution

The setup that follows incorporates all of the steps needed to determine the unknown lattice energy of MgF_2(s) from the given data.

Sublimation:	$\mathrm{Mg(s)} \longrightarrow \cancel{\mathrm{Mg(g)}}$	$\Delta H_1 = +146\ \mathrm{kJ}$
Bond dissociation:	$\mathrm{F_2(g)} \longrightarrow \cancel{\mathrm{2\,F(g)}}$	$\Delta H_2 = +159\ \mathrm{kJ}$
First ionization:	$\cancel{\mathrm{Mg(g)}} \longrightarrow \cancel{\mathrm{Mg^+(g)}} + \cancel{e^-}$	$\Delta H_3 = +738\ \mathrm{kJ}$
Second ionization:	$\cancel{\mathrm{Mg^+(g)}} \longrightarrow \cancel{\mathrm{Mg^{2+}(g)}} + \cancel{e^-}$	$\Delta H_4 = +1451\ \mathrm{kJ}$
Electron gain:	$\cancel{\mathrm{2\,F(g)}} + \cancel{2\,e^-} \longrightarrow \cancel{\mathrm{2\,F^-(g)}}$	$\Delta H_5 = 2(-328\ \mathrm{kJ})$
Unknown:	$\cancel{\mathrm{Mg^{2+}(g)}} + \cancel{\mathrm{2\,F^-}} \longrightarrow \mathrm{MgF_2(s)}$	$\Delta H_6 = \text{lattice energy}$
Overall:	$\mathrm{Mg(s)} + \mathrm{F_2(g)} \longrightarrow \mathrm{MgF_2(s)}$	$\Delta H_f^\circ = -1124\ \mathrm{kJ}$

The enthalpy of formation is equal to the sum of the energies of the individual processes.

$$\Delta H_f^\circ = -1124\ \mathrm{kJ} = (146 + 159 + 738 + 1451 - 656)\ \mathrm{kJ} + \text{lattice energy}$$

Rearranging the summed expression, we solve for the lattice energy.

$$\text{Lattice energy} = (-1124 - 146 - 159 - 738 - 1451 + 656)\ \mathrm{kJ} = -2962\ \mathrm{kJ}$$

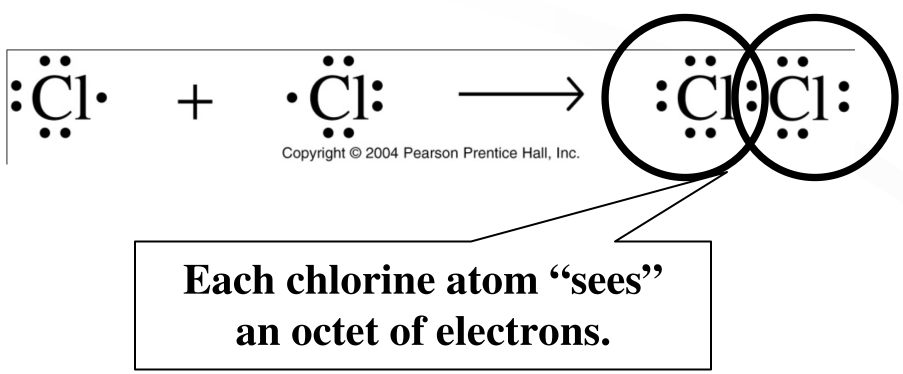
The lattice energy is -2962 kJ/mol MgF_2(s).

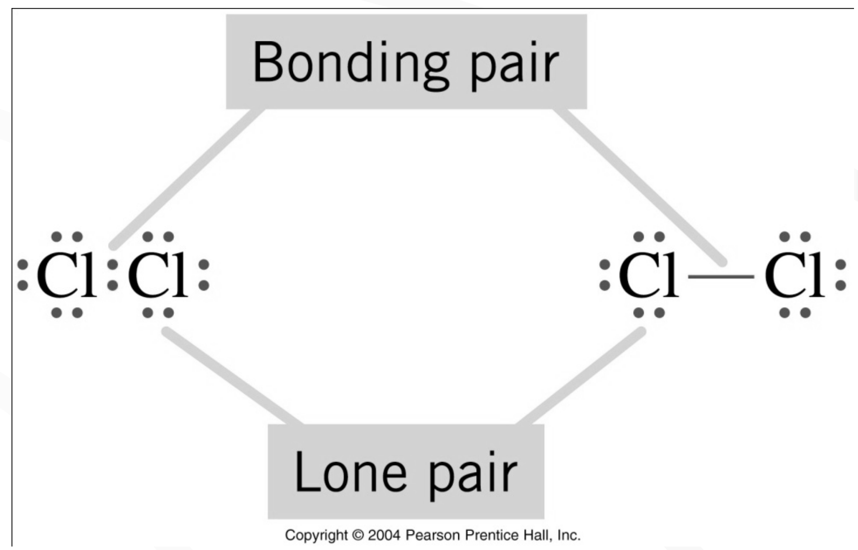
Lewis structures of simple molecules

- A ***Lewis structure*** is a combination of Lewis symbols that represents the formation of covalent bonds between atoms.
- In most cases, a Lewis structure shows the bonded atoms with the electron configuration of a noble gas; that is, the atoms obey the *octet rule*. (H obeys the *duet rule*.)
- The shared electrons can be counted for each atom that shares them, so each atom may have a noble gas configuration.

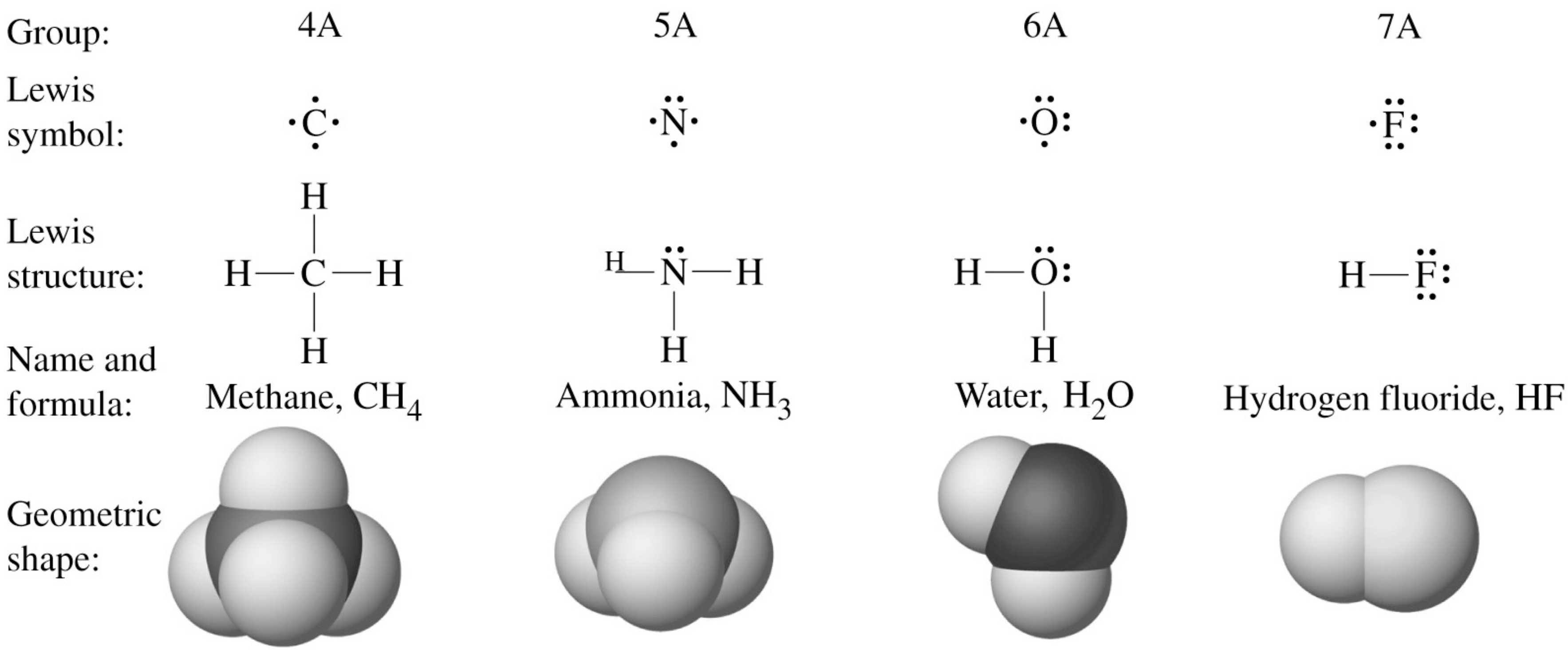
Lewis structures (cont'd)

- The shared pairs of electrons in a molecule are called ***bonding pairs***.
- The bonding pair is represented by a dash (—).
- The other electron pairs, which are not shared, are called ***nonbonding pairs***, or **lone pairs**.
- Two electrons from one element form a **coordinate covalent** bond (like in H_3O^+).





Some illustrative compounds

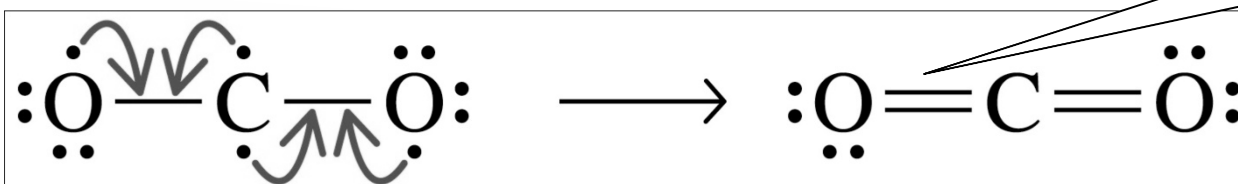


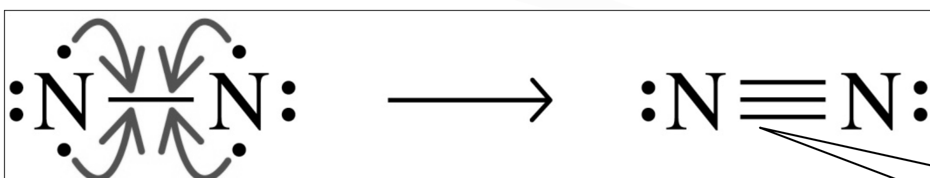
- Note that the two-dimensional Lewis structures do not necessarily show the correct *shapes* of the three-dimensional molecules. Nor are they intended to do so.
- The Lewis structure for water may be drawn with all three atoms in a line: H–O–H.
- We will learn how to predict *shapes* of molecules in Chapter 10.

Multiple covalent bonds

- The covalent bond in which one pair of electrons is shared is called a ***single bond***.
- Multiple bonds can also form:

In a *double bond* two pairs of electrons are shared.





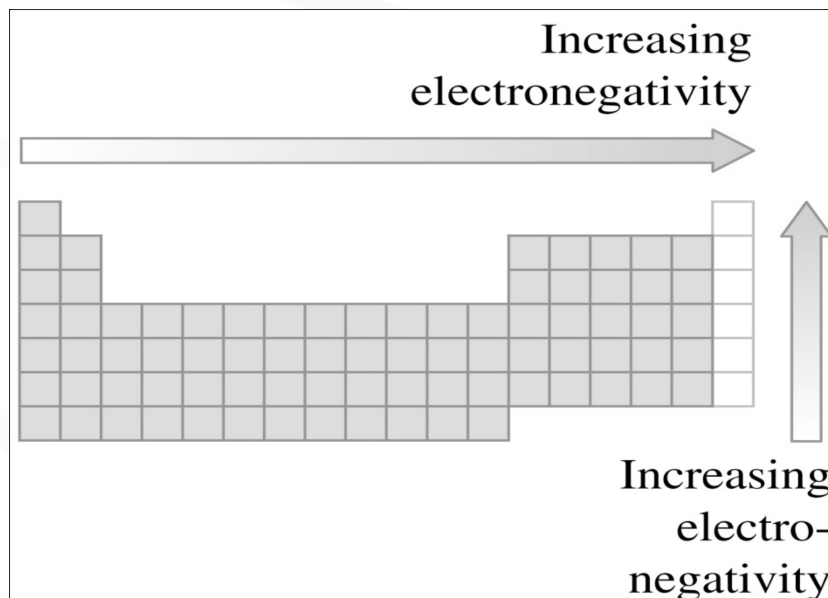
In a *triple bond* three pairs of electrons are shared.

Note that each atom obeys the octet rule, even with multiple bonds.
(Again, movement of a single electron <u>should</u> be represented by a single-headed arrow.)

Polar covalent bonds and electronegativity

- Sharing is not always equal
- **ELECTRONEGATIVITY** (**EN**) is a measure of the ability of an atom to attract its bonding electrons to itself.
- EN is related to ionization energy and electron affinity.
- The greater the EN of an atom in a molecule, the more strongly the atom attracts the electrons in a covalent bond.

Electronegativity generally ***increases*** from left to right within a period, and it generally ***increases*** from the bottom to the top within a group.



Pauling's electronegativities

F is the most electronegative, Fr is the least...
(It wouldn't hurt to remember the four elements of highest electronegativity: N, O, F, and Cl.)

- Below 1.0
- 1.0–1.4
- 1.5–1.9
- 2.0–2.4
- 2.5–2.9
- 3.0–4.0

Period	1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A
1	**H** 2.1																
2	**Li** 1.0	**Be** 1.5											**B** 2.0	**C** 2.5	**N** 3.0	**O** 3.5	**F** 4.0
3	**Na** 0.9	**Mg** 1.2											**Al** 1.5	**Si** 1.8	**P** 2.1	**S** 2.5	**Cl** 3.0
4	**K** 0.8	**Ca** 1.0	**Sc** 1.3	**Ti** 1.5	**V** 1.6	**Cr** 1.6	**Mn** 1.5	**Fe** 1.8	**Co** 1.8	**Ni** 1.8	**Cu** 1.9	**Zn** 1.7	**Ga** 1.6	**Ge** 1.8	**As** 2.0	**Se** 2.4	**Br** 2.8
5	**Rb** 0.8	**Sr** 1.0	**Y** 1.2	**Zr** 1.4	**Nb** 1.6	**Mo** 1.8	**Tc** 1.9	**Ru** 2.2	**Rh** 2.2	**Pd** 2.2	**Ag** 1.9	**Cd** 1.7	**In** 1.7	**Sn** 1.8	**Sb** 1.9	**Te** 2.1	**I** 2.5
6	**Cs** 0.7	**Ba** 0.9	**La*** 1.1	**Hf** 1.3	**Ta** 1.5	**W** 1.7	**Re** 1.9	**Os** 2.2	**Ir** 2.2	**Pt** 2.2	**Au** 2.4	**Hg** 1.9	**Tl** 1.8	**Pb** 1.8	**Bi** 1.9	**Po** 2.0	**At** 2.2
7	**Fr** 0.7	**Ra** 0.9	**Ac†** 1.1														

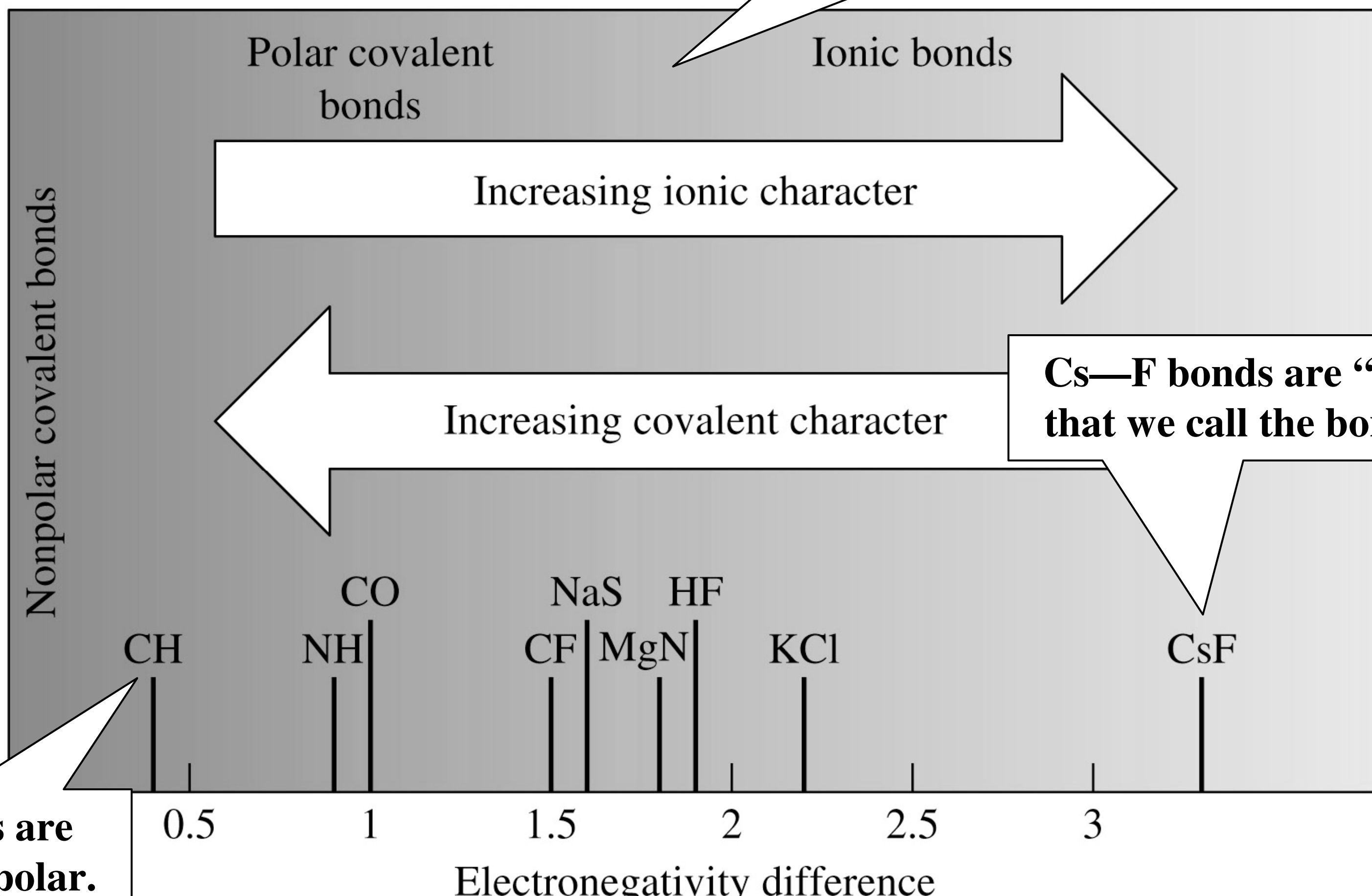
*Lanthanides: 1.1–1.3
†Actinides: 1.3–1.5

Electronegativity (cont'd)

- The greater the difference in electronegativity the greater the *ionic character*...the more ionic character to the bond, the more *polar* it is.
- A homonuclear diatomic molecule is completely *nonpolar*.
- Polarity is drawn using an arrow with a cross through the tail (3 slides ahead).
- "Partially positive" atoms are denoted by a "delta-plus“ (δ+), partially negative atoms are denoted by a "delta-minus“(δ-).

Electronegativity difference and bond type

No sharp cutoff between ionic and covalent bonds, though some books say 1.7.



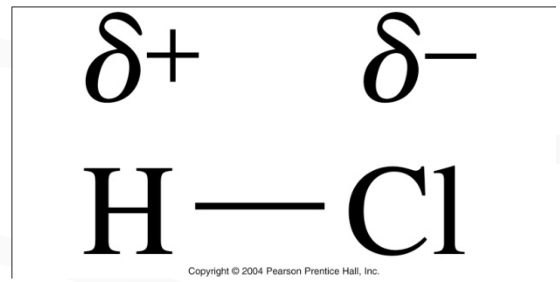
Cs—F bonds are "so polar" that we call the bonds ____.

C—H bonds are virtually nonpolar.

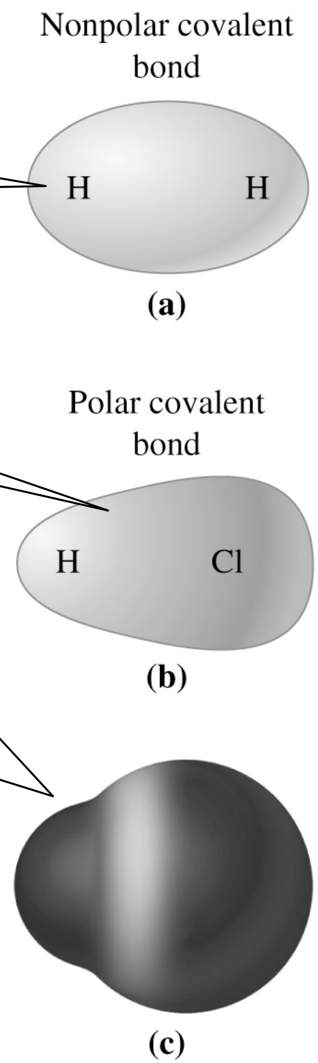
Depicting polar covalent bonds

In nonpolar bonds, electrons are shared *equally*.

Polar bonds are also depicted by partial positive and partial negative symbols …

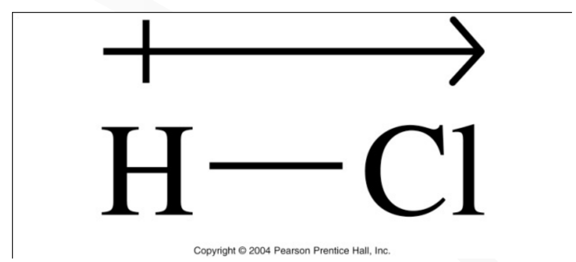


***Unequal* sharing in polar covalent bonds.**



Polar bonds are often depicted using colors to show electrostatic potential (blue = positive, red = negative).

… or with a cross-based arrow pointing to the more electronegative element.



Example

- Use electronegativity values to arrange the following bonds in order of increasing polarity:

  Br—Cl, Cl—Cl, Cl—F, H—Cl, I—Cl

- Worked out next slide

Example worked out

Strategy

We begin with the knowledge that the polar character of a bond is determined by electronegativity differences. For the individual electronegativities of the atoms, we turn to Figure 9.8.

Solution

The electronegativities (EN) and electronegativity differences are

EN:	2.8	3.0	3.0	3.0	3.0	4.0	2.1	3.0	2.5	3.0
	Br — Cl		Cl — Cl		Cl — F		H — Cl		I — Cl	
ΔEN:	0.2		0.0		1.0		0.9		0.5	

The order of increasing polarity is therefore

Cl — Cl < Br — Cl < I — Cl < H — Cl < Cl — F.

- **Exercise 9.5A**

  Use electronegativity values to arrange the following bonds in order of increasing polarity:
  C — Cl, C — H, C — Mg, C — O, C — S

- **Exercise 9.5B**

  (a) Of the elements carbon, silicon, bromine, sulfur, nitrogen, chlorine, arsenic, fluorine, and selenium, which two form the covalent bond that has the most ionic character?
  (b) Which two form the covalent bond that has the least ionic character?

Writing Lewis structures: skeletal structures

- The SKELETAL STRUCTURE shows the arrangement of atoms.
- Lewis structures have terminal atoms drawn around a central atom....think of each atom as a four-sided box that must obey the octet (duet) rule.....draw dashes for single bonds, then add *lone pairs* of electrons to the terminal atoms to get an octet...then add lone pairs and/or form multiple bonds to central atoms as needed to account for the *total number of valence electrons*
- Hydrogen atoms are ***terminal atoms*** (bonded to only one other atom).
- The ***central atom*** of a structure usually has the ***lowest*** electronegativity.
- In oxoacids ($HClO_4$, HNO_3, etc.) hydrogen atoms are usually bonded to ***oxygen*** atoms.
- Molecules and polyatomic ions usually have compact, symmetrical structures.
- C, N, O, and S are often double bonded. C and N can be triple bonded.

Example

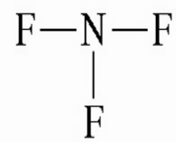
– Write the Lewis structure of nitrogen trifluoride, NF_3.

Solution

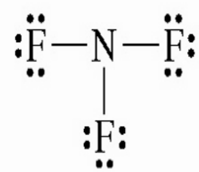
Step 1: *Determine the number of valence electrons.*

In one N atom (group 5A) and three F atoms (group 7A), there are $5 + (3 \times 7) = 26$ valence electrons.

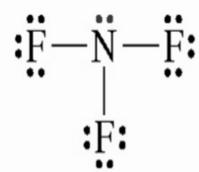
Step 2: *Write a skeletal structure.* The electronegativity of N is 3.0; that of F is 4.0. We expect a skeletal structure with N as a central atom and F as terminal atoms. The three nitrogen-to-fluorine bonds in this structure account for six electrons.



Step 3: *Complete the octets of terminals atoms.* We complete the octets of the F atoms by placing three lone pairs of electrons around each. This accounts for 18 additional electrons.

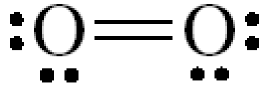


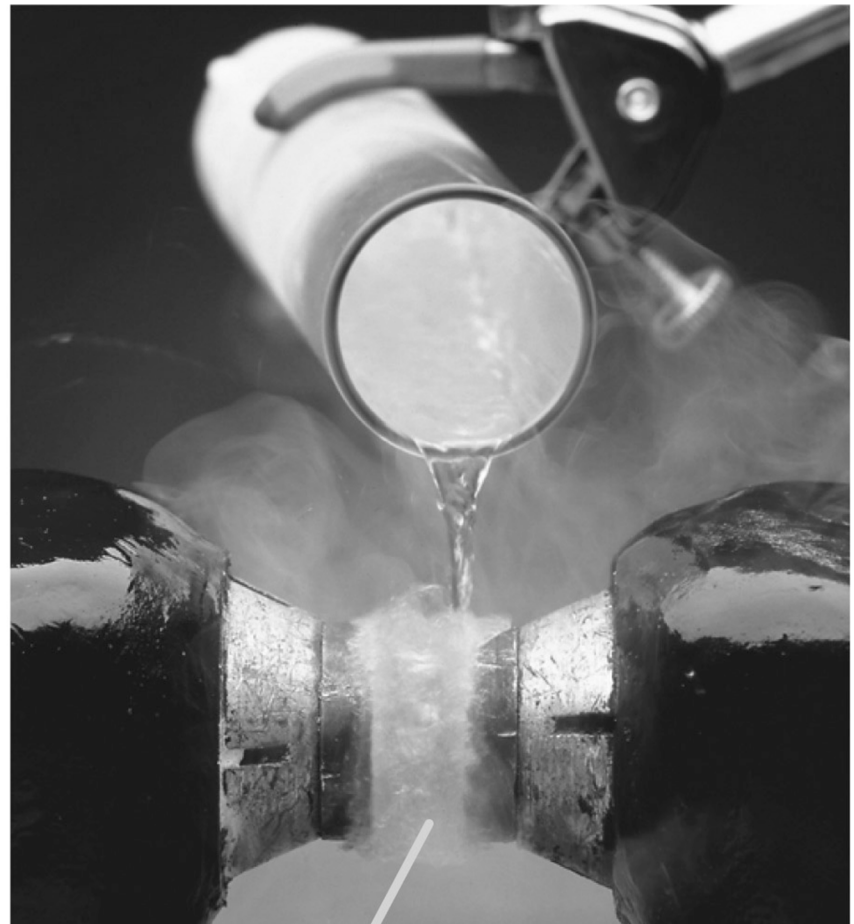
Step 4: *Assign lone pairs to central atom(s).* We have now assigned $6 + 18 = 24$ of the 26 valence electrons. We place the remaining two as a lone pair on the N atom.



- Write the Lewis structure of hydrazine, N_2H_4.

Lewis structures are not always "right"–
The importance of experimental evidence

- The Lewis structure commonly drawn for oxygen is 
- But oxygen is ***paramagnetic***, and therefore must have ***unpaired*** electrons.
- Lewis structures are a useful ***tool***, but they do not always represent molecules correctly, even when the Lewis structure is plausible.



Liquid oxygen is paramagnetic and is held between the poles of the magnet.

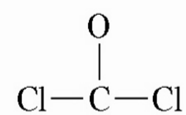
Example

Write a plausible Lewis structure for phosgene, $COCl_2$.

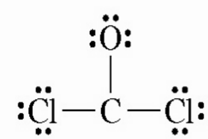
Step 1: *Determine the number of valence electrons.*

In one C atom (group 4A), one O atom (group 6A), and two Cl atoms (group 7A), there are $4 + 6 + (2 \times 7) = 24$ valence electrons.

Step 2: *Write a skeletal structure.* The electronegativities are 2.5 for C, 3.5 for O, and 3.0 for Cl. We expect C, the least electronegative atom, to be the central atom, and O and Cl to be terminal atoms attached to it. This skeletal structure accounts for six of the valence electrons.

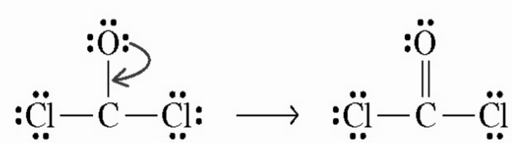


Step 3: *Complete the octets of terminal atoms.* Place three lone pairs around the O atom and three lone pairs around each of the Cl atoms, for a total of 18 electrons. This completes the octets of the terminal atoms.



Step 4: *Assign lone pairs to central atom(s).* Up to this point, we have assigned $6 + 18 = 24$ of the 24 valence electrons, so there are none available to complete the octet of the central C atom. Therefore we must go to step 5.

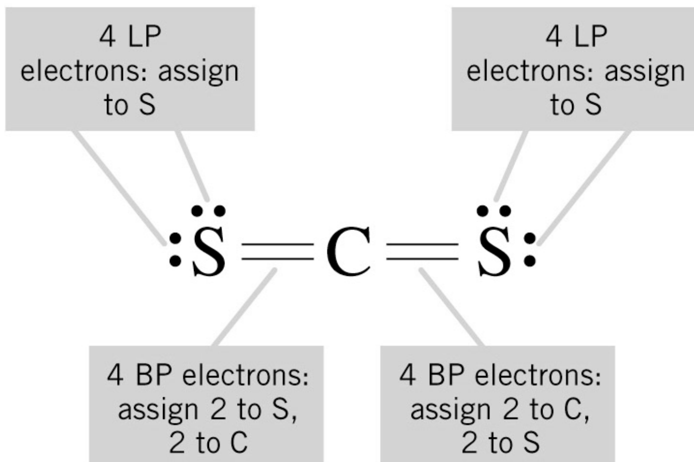
Step 5: *Form multiple bonds to complete octets of central atom(s).* Complete the octet on the C atom by shifting a lone pair of electrons from the O atom to form a carbon-to-oxygen double bond. (C and O are two atoms that are able to form a double bond.)



Formal charge

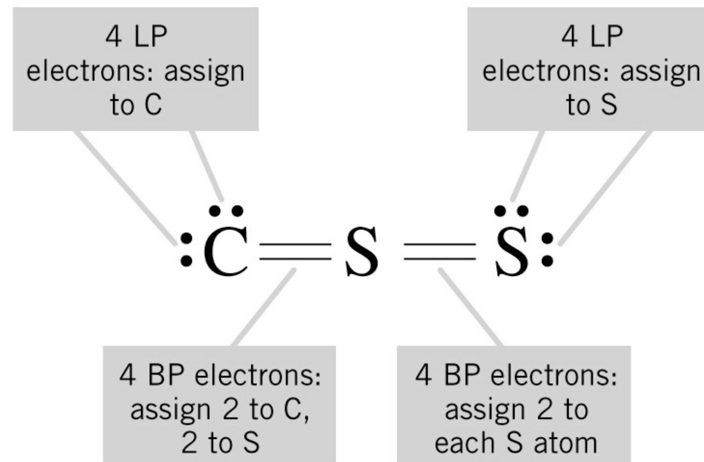
- ***Formal charge*** is the difference between the number of valence electrons in a free (uncombined) atom and the number of electrons assigned to that atom when bonded to other atoms in a Lewis structure.
- Formal charge is a ***hypothetical*** quantity; a useful tool in predicting reactivity.
- Usually, the ***most plausible*** Lewis structure is one with no formal charges.
- When formal charges are required, they should be as small as possible.
- Negative formal charges should appear on the most electronegative atoms. *Makes sense, right?*
- Adjacent atoms in a structure should *not* carry formal charges of the same sign.

Formal charge illustrated



	S	C	S
Valence electrons	6	4	6
Electrons assigned	6	4	6
Formal charge	0	0	0

(a)



	C	S	S
Valence electrons	4	6	6
Electrons assigned	6	4	6
Formal charge	−2	+2	0

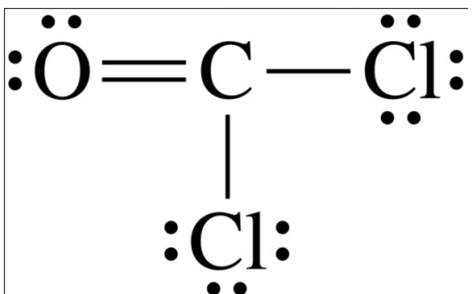
(b)

✓ yes

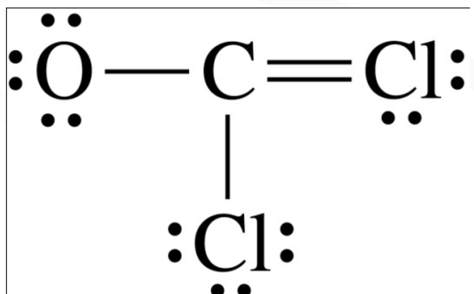
✗ no

Example

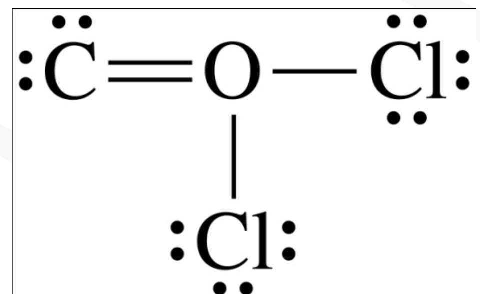
- Before, we wrote a Lewis structure for the molecule $COCl_2$, shown here as structure (a). Show that structure (a) is more plausible than (b) or (c).



(a)



(b)

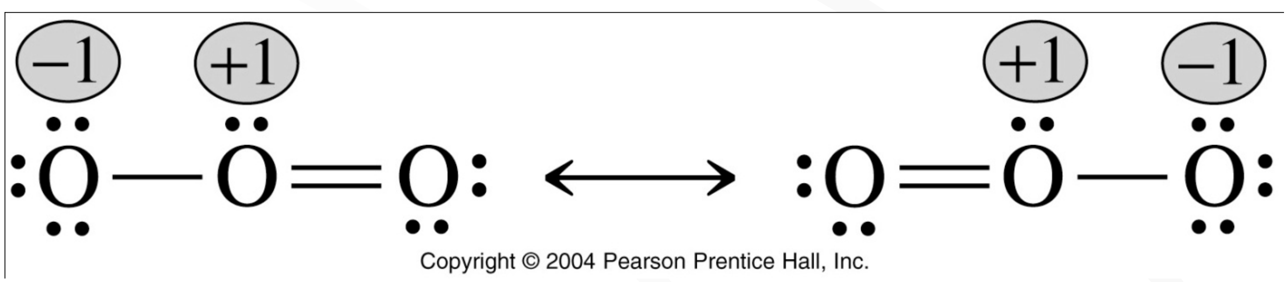


(c)

Resonance: delocalized bonding

- When a molecule or ion can be represented by two or more plausible Lewis structures that differ ***only in the distribution of electrons***, the true structure is a composite, a hybrid, of them.
- The different plausible structures are called ***resonance structures***.
- The actual molecule or ion that is a hybrid of the resonance structures is called a ***resonance hybrid***.
- Electrons that are part of the resonance hybrid are spread out over several atoms and are referred to as being ***delocalized***.

***Three* pairs of electrons are distributed among *two* bonds.**



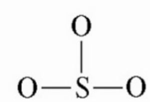
Example

- Write three equivalent Lewis structures for the SO_3 molecule that conform to the octet rule, and describe how the resonance hybrid is related to the three structures.

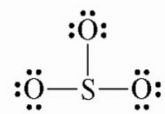
Step 1: *Determine the number of valence electrons.*

There are $6 + (3 \times 6) = 24$ valence electrons.

Step 2: *Write a skeletal structure.* The skeletal structure has sulfur, the atom of lower electronegativity, as the central atom.

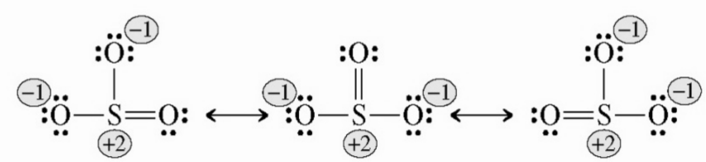


Step 3: *Complete the octets of terminal atoms.* Place three lone pairs of electrons on each O atom to complete its octet.



Step 4: *Assign lone pairs to central atom(s).* We have already assigned all 24 valence electrons; there are none left to place as lone pairs on the central atom. Because the S atom does not yet have an octet, we must go on to step 5.

Step 5: *Form multiple bonds to complete octets of central atom(s).* Move a lone pair of electrons from any terminal O atom to form a double bond to the central S atom. Because the double bond can go to any one of the three terminal O atoms, we get three equivalent structures that differ only in the position of the double bond.



Resonance structures

Molecules that do not follow the octet rule

- Molecules with an ***odd number*** of valence electrons have at least one of them unpaired and are called ***free radicals***.
- Some molecules have ***incomplete octets***. These are usually compounds of Be, B, or Al; they generally have some unusual bonding characteristics, and are often quite reactive.
- Some compounds have ***expanded valence shells***, which means that the central atom has more than eight electrons around it.
- A central atom can have expanded valence if it is in the third period or lower (i.e., S, Cl, P).

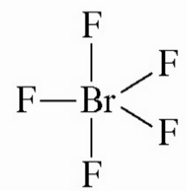
Example

- Write the Lewis structure for bromine pentafluoride, BrF_5.

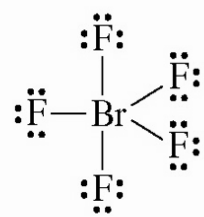
Step 1: *Determine the number of valence electrons.*

Both Br and F are in group 7A. All the atoms in the structure have seven valence electrons, and $7 + (5 \times 7) = 42$ valence electrons must appear in the Lewis structure.

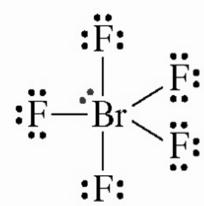
Step 2: *Write a skeletal structure.* The skeletal structure has Br, the atom of lower electronegativity, as the central atom.



Step 3: *Complete the octets of terminal atoms.* To complete the octets of the F atoms, we place three lone pairs of electrons around each one.

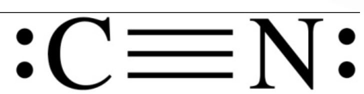


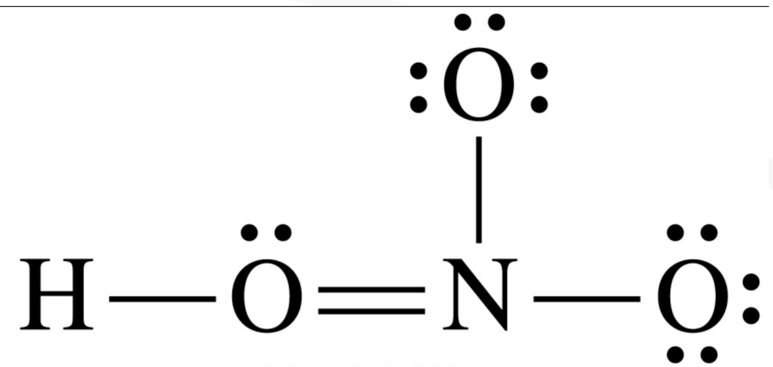
Step 4: *Assign lone pairs to central atom(s).* Through step 3, we have assigned 40 of the 42 valence electrons; two remain to be placed. They can be shown as a lone pair on the bromine atom. In this representation, the Br atom has an expanded valence shell containing 12 electrons.



Example

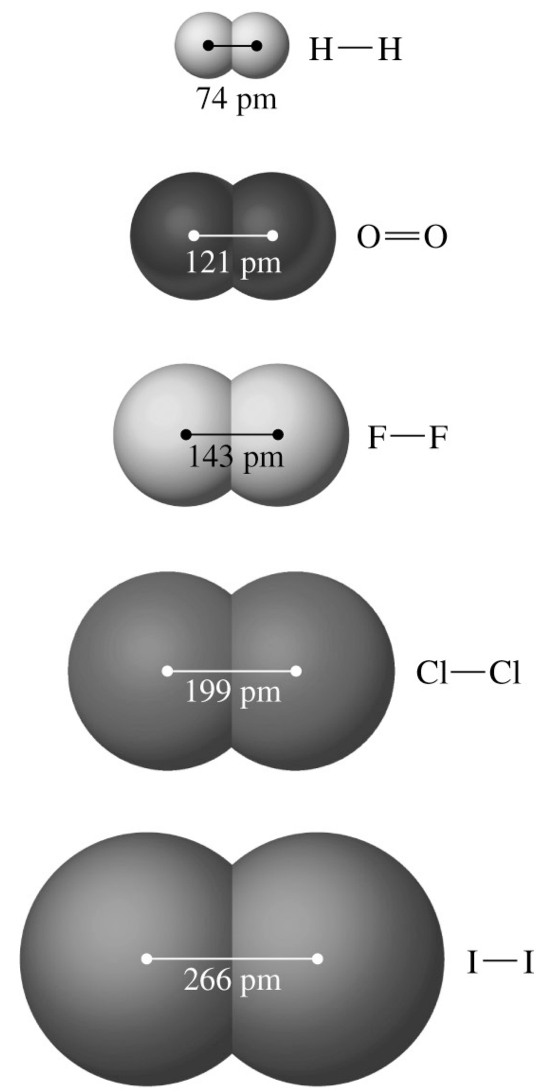
- Indicate the error in each of the following Lewis structures. Replace each by a more acceptable structure(s).

a. 

b. 

Bond order and bond length

- ***Bond order*** is the number of shared electron pairs in a bond.
- A single bond has BO = 1, a double bond has BO = 2, etc.
- ***Bond length*** is the distance between the nuclei of two atoms joined by a covalent bond.
- Bond length depends on the particular atoms in the bond and on the bond order.



Bond energy

- ***Bond-dissociation energy*** (***D***) is the energy required to break one mole of a particular type of covalent bond in a *gas*-phase compound.
- Energies of some bonds can differ from compound to compound, so we use an ***average*** bond energy.

The H—H bond energy is precisely known …

… while the O—H bond energies for the two bonds in H_2O are different.

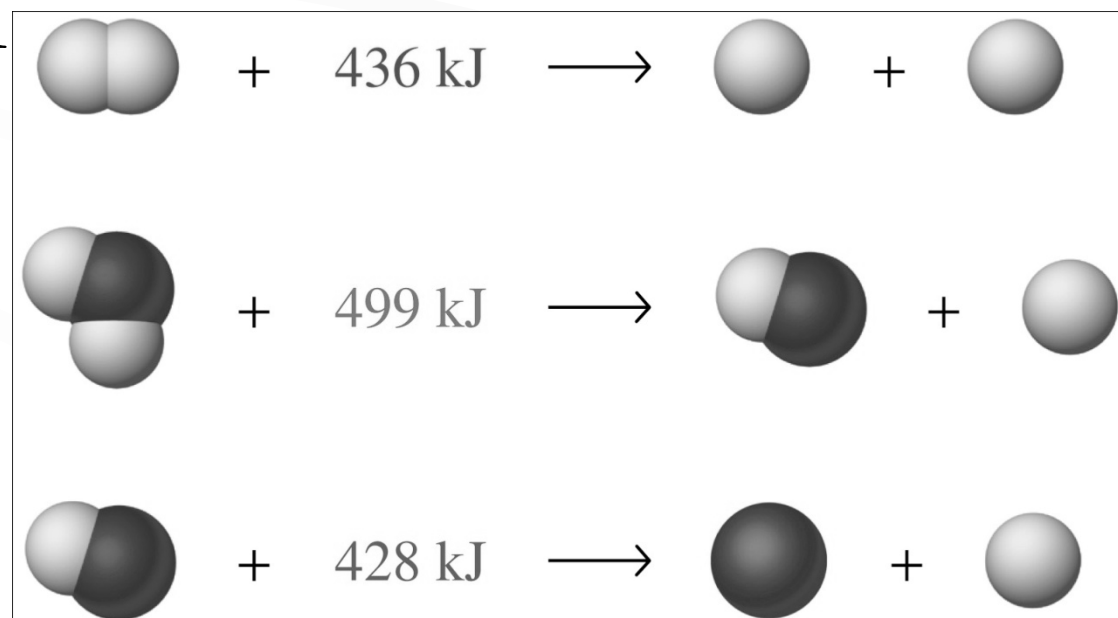


Table 9.1 Some Representative Bond Lengths and Bond Energies

Bond	Bond Length, pm	Bond Energy,[a] kJ/mol	Bond	Bond Length, pm	Bond Energy,[a] kJ/mol
H—H	74	436	C—O	143	360
H—C	110	414	C═O	120	736[b]
H—N	100	389	C—Cl	178	339
H—O	97	464	N—N	145	163
H—S	132	368	N═N	123	418
H—F	92	565	N≡N	110	946
H—Cl	127	431	N—O	136	222
H—Br	141	364	N═O	120	590
H—I	161	297	O—O	145	142
C—C	154	347	O═O	121	498
C═C	134	611	F—F	143	159
C≡C	120	837	Cl—Cl	199	243
C—N	147	305	Br—Br	228	193
C═N	128	615	I—I	266	151
C≡N	116	891			

[a] Bond-dissociation energy for the bonds in diatomic molecules (H_2, HF, HCl, HBr, HI, N_2, O_2, F_2, Cl_2, Br_2 and I_2) and average bond energies for the other bonds.

[b] The value for the CO bond in CO_2 is considerably different: 799 kJ/mol.

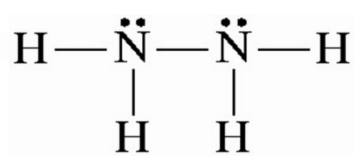
Trends in bond lengths and energies

- The *higher* the order (for a particular type of bond), the ***shorter*** and the ***stronger*** (higher energy) the bond.
- A N=N *double* bond is shorter and stronger than a N–N *single* bond.
- There are *four* electrons between the two positive nuclei in N=N. This produces more electrostatic attraction than the two electrons between the nuclei in N–N.

Example

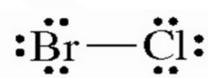
- Estimate the length of
  (a) the nitrogen-to-nitrogen bond in N_2H_4 and
  (b) the bond in BrCl.

(a) The Lewis structure



is a plausible one: valence-shell octets for the N atoms, duets for the H atoms, and formal charges of 0. The nitrogen-to-nitrogen bond is single, and from Table 9.1 we find its bond length to be 145 pm.

(b) To get the Lewis structure of BrCl, we imagine substituting one Br atom for one Cl atom in the Lewis structure of Cl_2, arriving at the structure



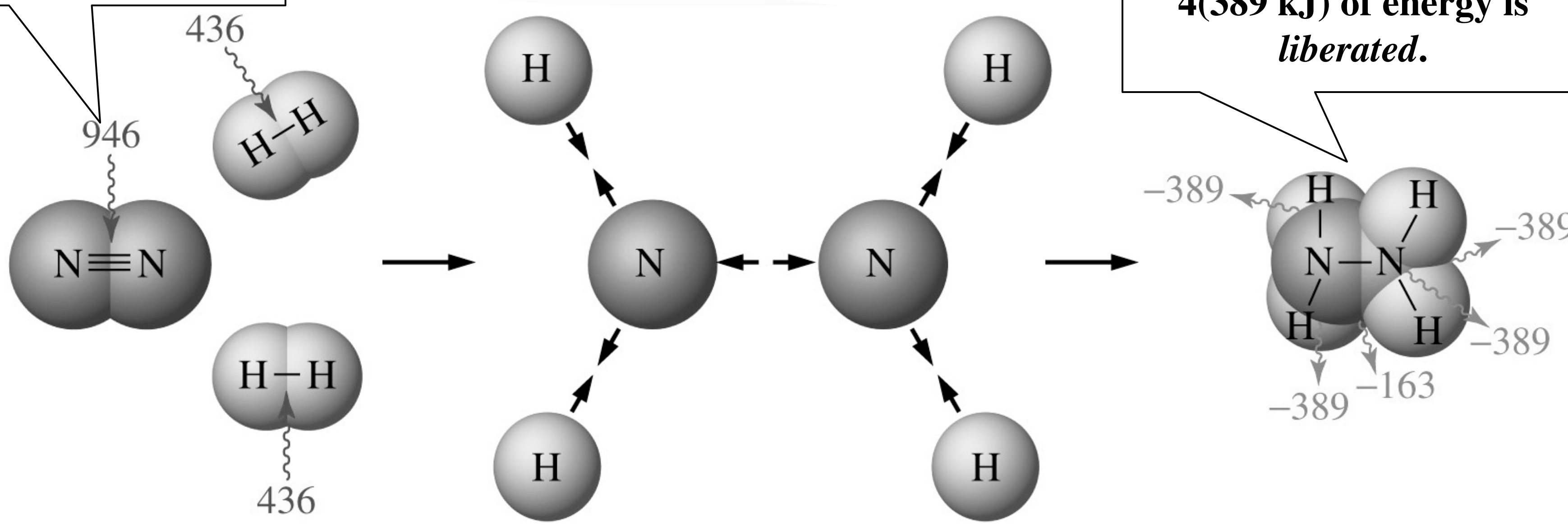
The BrCl molecule contains a Br—Cl single bond with a length approximately one-half the Cl—Cl bond length *plus* one-half the Br—Br bond length. From Table 9.1, we get $\left[\left(\frac{1}{2} \times 199\right) + \left(\frac{1}{2} \times 228\right)\right] = 214$ pm.

Calculations involving bond energies

For the reaction $N_2(g) + 2\,H_2(g) \rightarrow N_2H_4(g)$ to occur …

… we must *supply* 946 kJ …

When the bonds of the product *form*, 163 kJ plus 4(389 kJ) of energy is *liberated*.



… plus 2(436 kJ), to *break* bonds of the reactants.

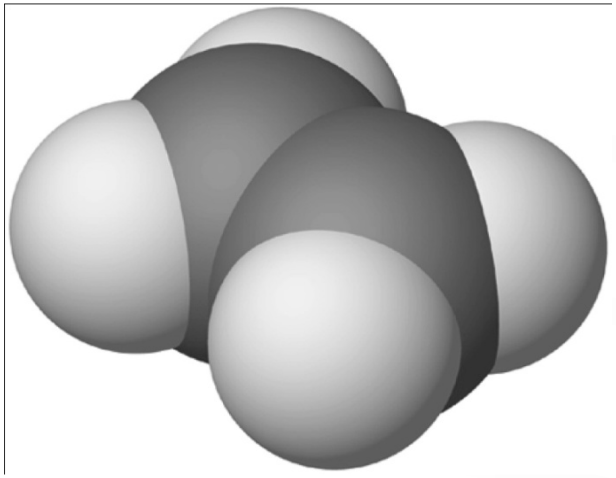
$$\Delta H = (+946\text{ kJ}) + 2(+436\text{ kJ}) + (-163\text{ kJ}) + 4(-389\text{ kJ})$$

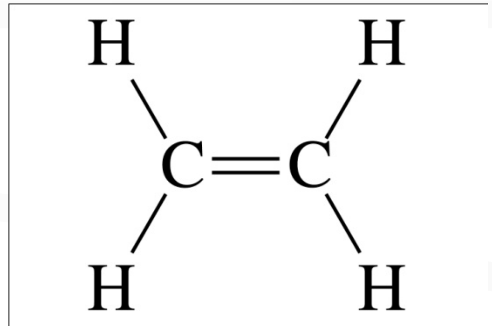
Flashback: alkenes and alkynes

- Hydrocarbons with double or triple bonds between carbon atoms are called ***unsaturated hydrocarbons***.
- ***Alkenes*** are hydrocarbons with one or more C=C double bonds.
- The simplest alkene is C_2H_4, ethene (ethylene).
- ***Alkynes*** are hydrocarbons that have one or more carbon–carbon triple bonds.
- The simplest alkyne is C_2H_2, ethyne (acetylene).

Molecular models of ethene and ethyne

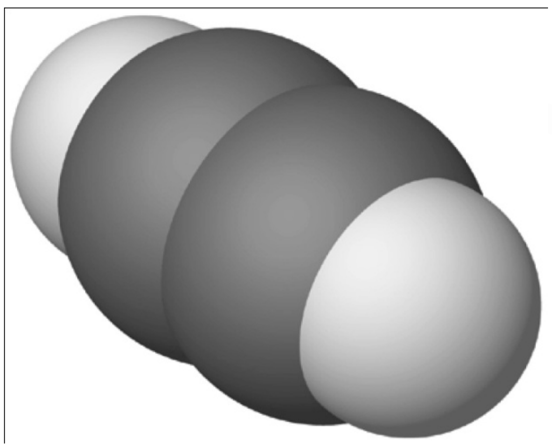
space-filling models





Structural formula

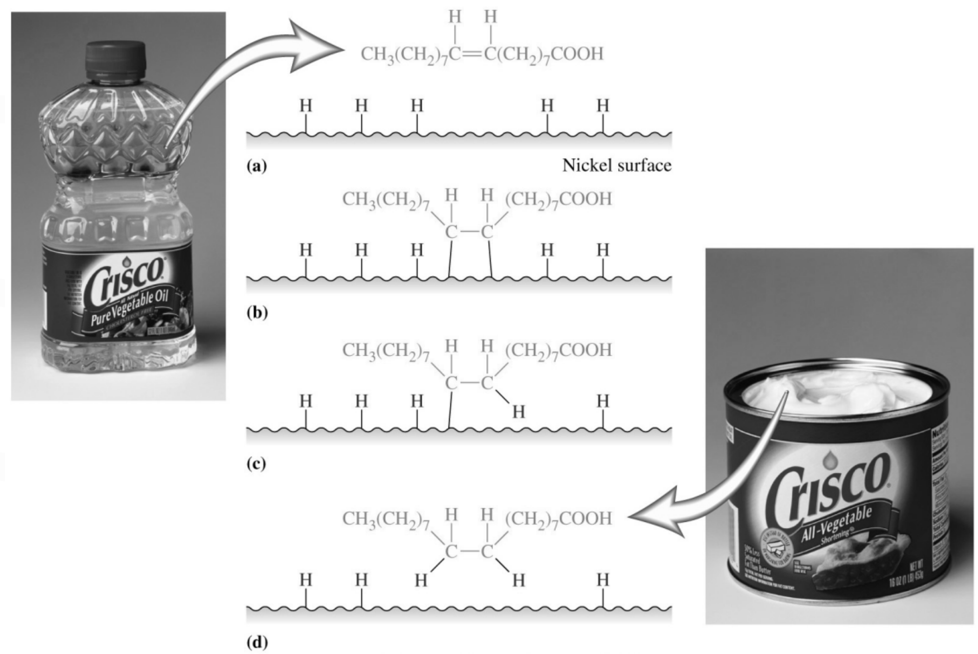
$$\underbrace{CH_2{=}CH_2 \quad \text{or} \quad H_2C{=}CH_2}_{\text{Condensed structural formula}}$$

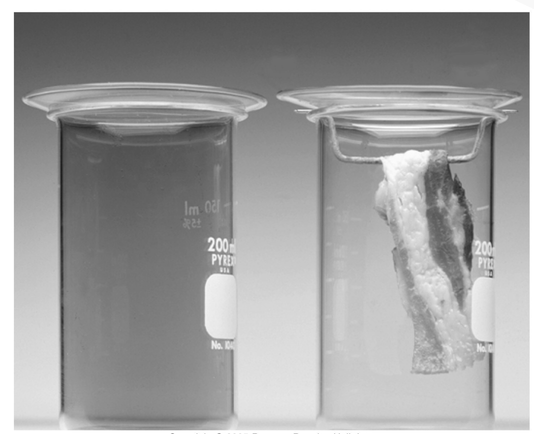


H:C:::C:H or $H{-}C{\equiv}C{-}H$

Fats and oils

- Triglycerides are composed of glycerol plus three long-chain carboxylic (fatty) acids
- What's the difference between a fat and an oil? Fats are solids at RT, Oils are liquids
- Unsaturated fats become saturated upon *hydrogenation* (Fig. 13.21)
- Your book also has a nice bromination picture in this chapter





Polymers

- ***Polymers*** are compounds in which many identical molecules have been joined together.
- ***Monomers*** are the simple molecules which join together to form polymers.
- Often, the monomers have double or triple bonds.
- The process of these molecules joining together is called ***polymerization***.
- Many everyday products and many biological compounds are polymers.

Formation of polyethylene

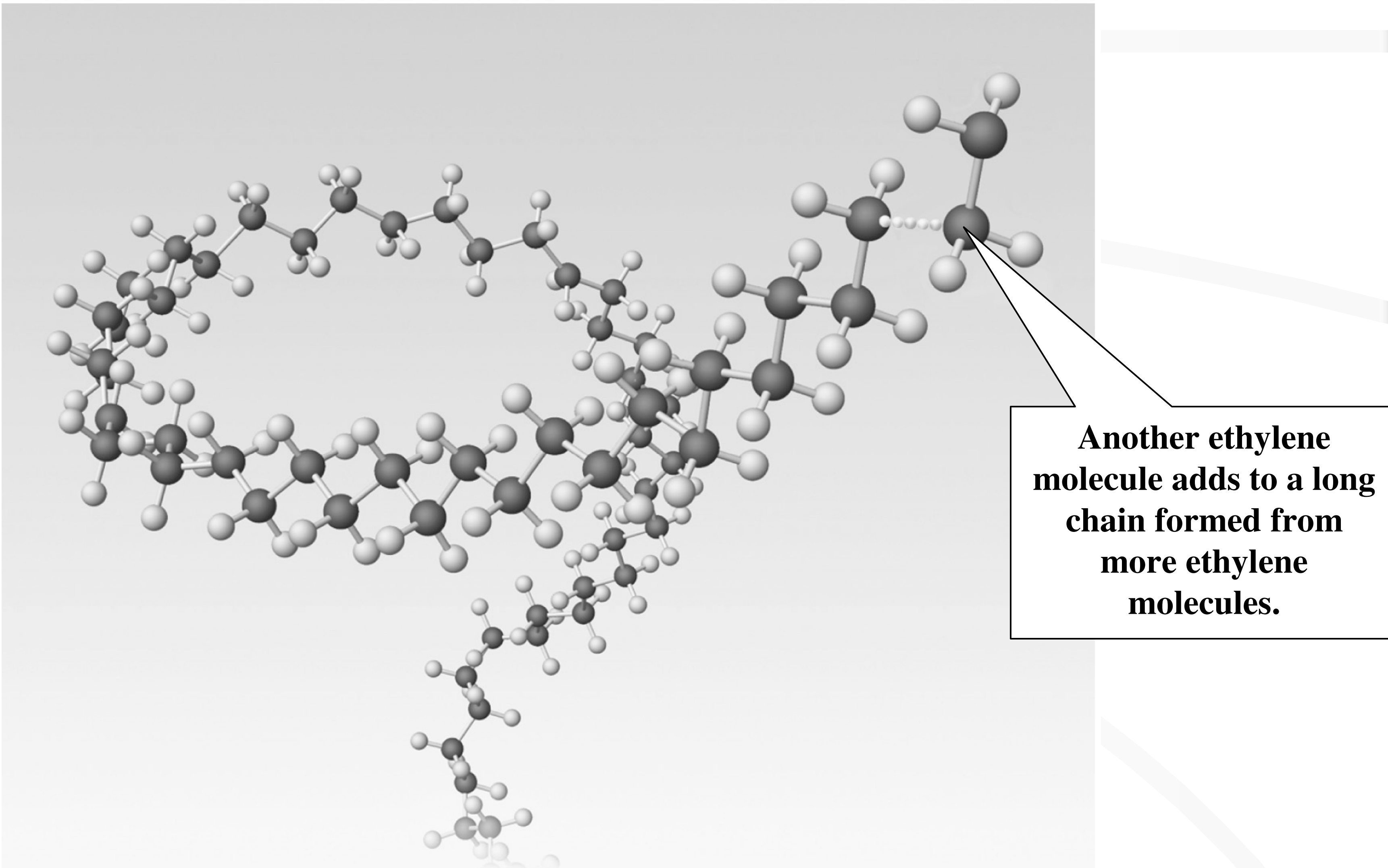
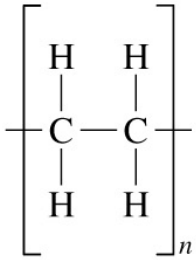
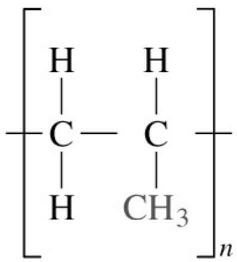
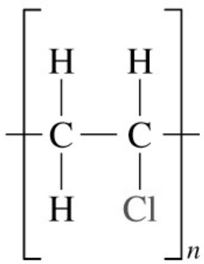
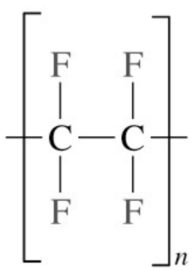


Table 9.2 A Selection of Addition Polymers

Monomer	Polymer	Polymer Name	Some Uses
$H_2C{=}CH_2$		Polyethylene	Plastic bags, milk bottles, toys, electrical insulation, plastic wrap, paper coatings, food containers and lids, storage containers, laboratory bottles, wastebaskets
$H_2C{=}CH{-}CH_3$		Polypropylene	Indoor–outdoor carpeting, bottles, insulating sportswear, tear-resistant envelopes, labels, disposable clothing and protective outerwear, moisture-resistant packaging, chemical-resistant plumbing pipe
$H_2C{=}CH{-}Cl$		Poly(vinyl chloride), PVC	Plastic wrap, simulated leather (Naugahyde), T-shirt printing, garden hoses, clear flexible tubing, "soft" toys, plumbing fixtures and pipe, vinyl flooring, rainwear
$F_2C{=}CF_2$		Polytetrafluoroethylene, Teflon	Nonstick coating for cooking utensils, electrical insulation, chemical-resistant laboratory ware, coatings for clothing and carpet

Plastics

- Although plastics are marked recyclable 1-7, most areas only recycle plastics 1 and 2
  - \- Polyethylene terephthalate (**PET**)
  - \- High-density polyethylene (**HDPE**)
  - \- Polyvinyl chloride (**PVC**)
  - \- Low-density polyethylene (**LDPE**)
  - \- Polypropylene (**PP**)
  - \- Polystyrene (**PS**)
  - **7** - Other resins, like acrylonitrile butadine styrene (**ABS**)