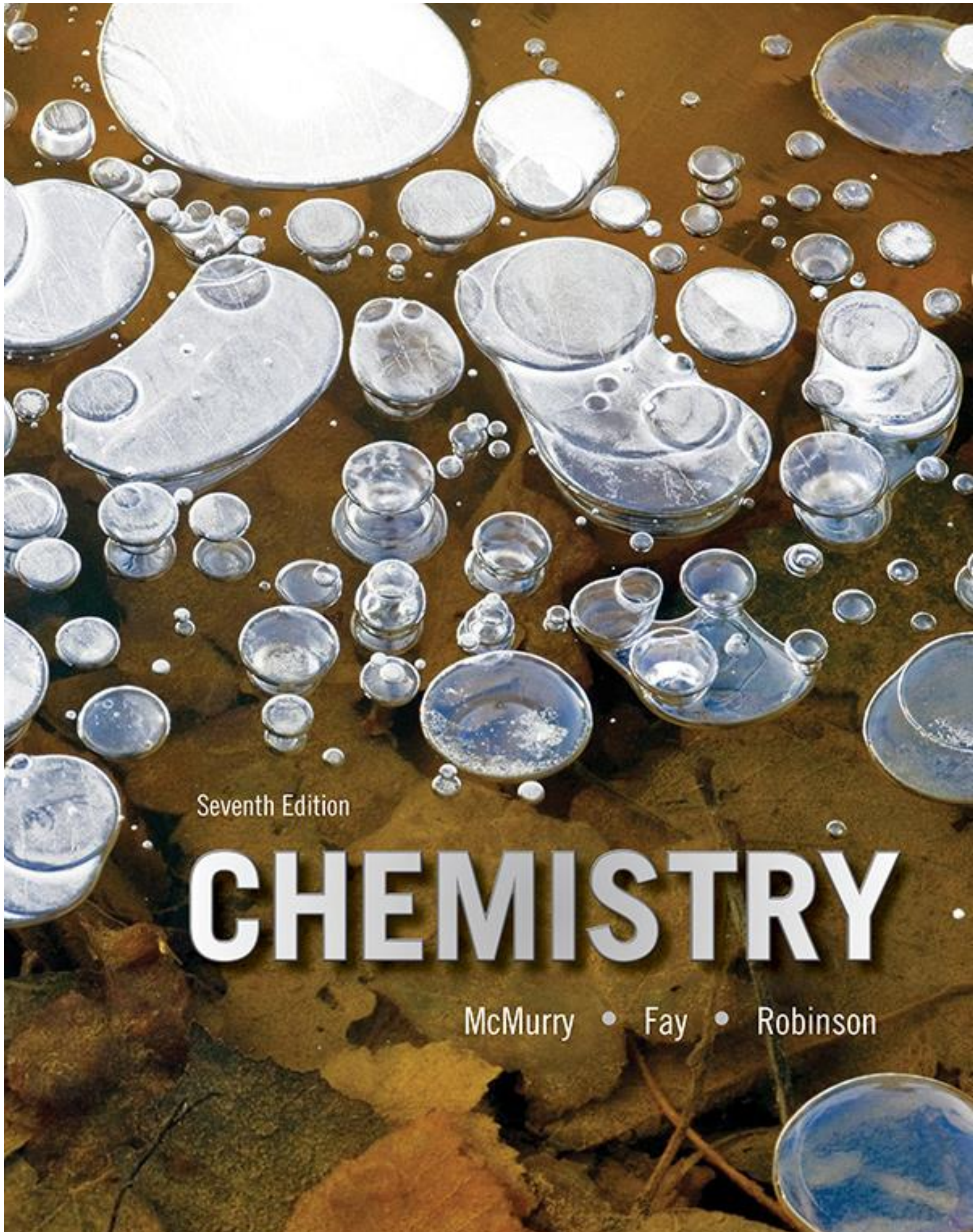


Lecture Presentation

Chapter 7

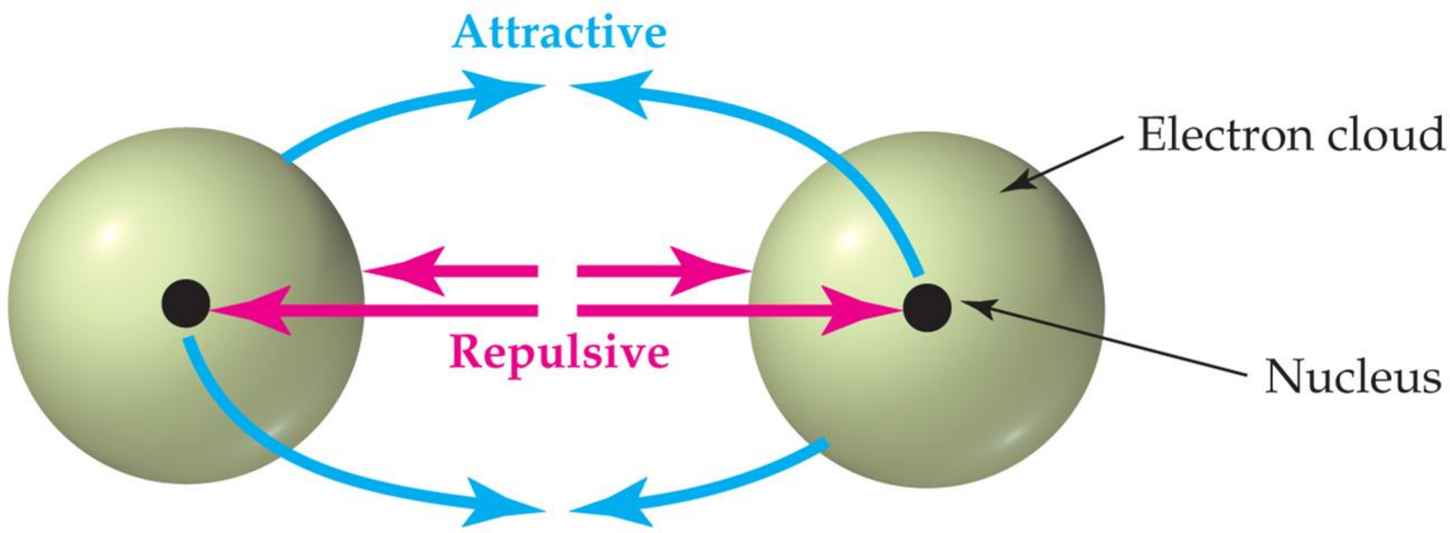
Covalent Bonding and Electron-Dot Structures

HW: 7.1, 7.2, 7.3, 7.4, 7.5, 7.6, 7.7, 7.8, 7.9, 7.10, 7.11, 7.13, 7.14, 7.15, 7.16, 7.18, 7.19, 7.20, 7.21, 7.22, 7.25, 7.34, 7.38, 7.40, 7.42, 7.44, 7.46, 7.48, 7.50, 7.54, 7.62, 7.80, 7.84, 7.86

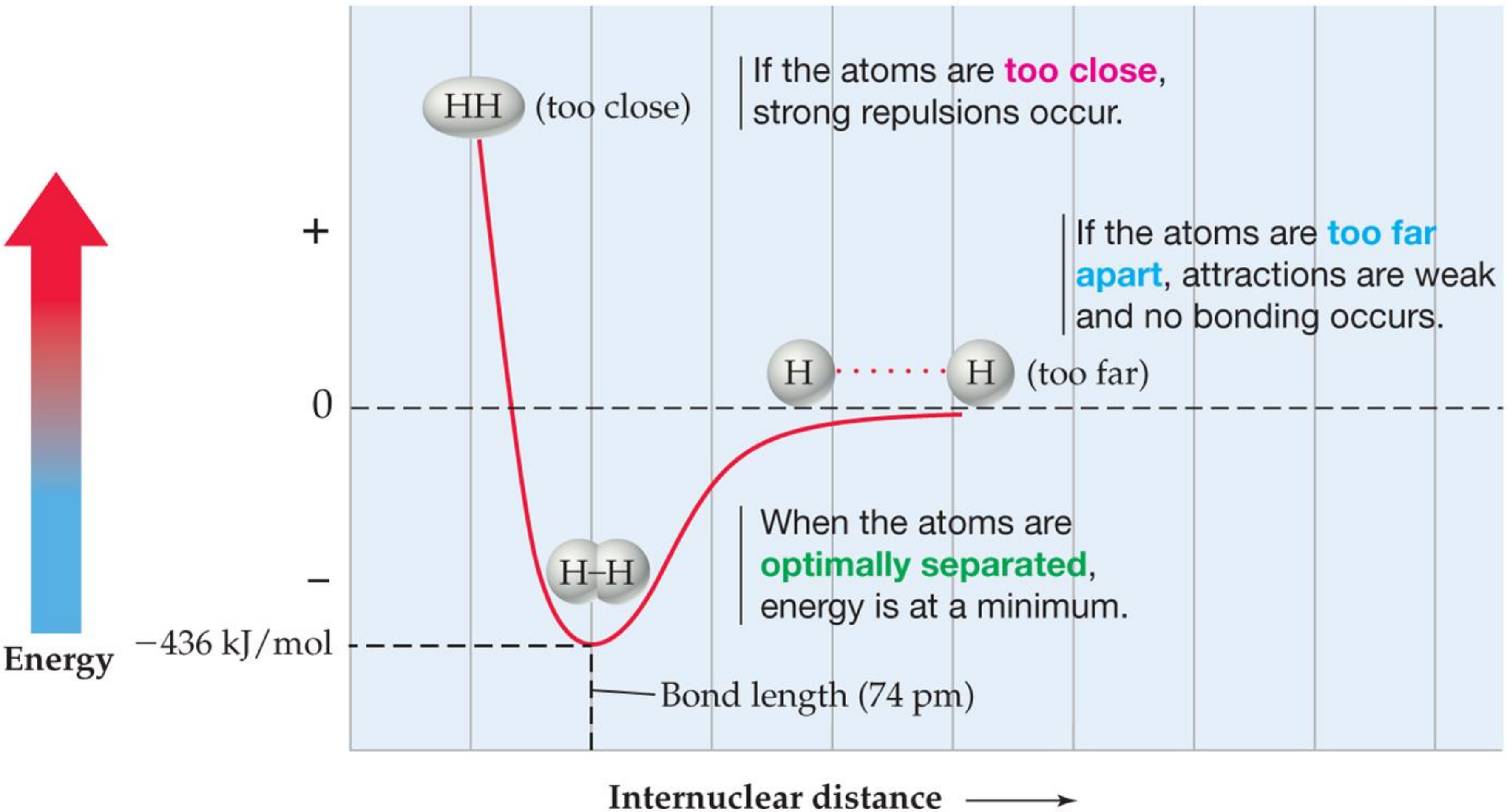
John E. McMurry
Robert C. Fay

Covalent Bonding in Molecules

Covalent Bond: A bond that results from the sharing of electrons between atoms



The nucleus–electron **attractions** are greater than the nucleus–nucleus and electron–electron **repulsions**, resulting in a net attractive force that binds the atoms together.



HH (too close)
If the atoms are too close, strong repulsions occur.
If the atoms are too far apart, attractions are weak and no bonding occurs.
H H (too far)
When the atoms are optimally separated, energy is at a minimum.
H–H
+
0
–
−436 kJ/mol
Energy
Bond length (74 pm)
Internuclear distance

Strengths of Covalent Bonds (don't memorize numbers) (shorter bonds are stronger bonds) (bond length also related to atom radius) (multiple bonds are shorter & stronger*)

TABLE 7.1 Average Bond Lengths (pm)

H—H	74[a]	C—H	110	N—H	98	O—F	130	I—I	267[a]
H—C	110	C—C	154	N—C	147	O—Cl	165	S—F	168
H—F	92[a]	C—F	141	N—F	134	O—Br	180	S—Cl	203
H—Cl	127[a]	C—Cl	176	N—Cl	169	O—I	199	S—Br	218
H—Br	142[a]	C—Br	191	N—Br	184	O—N	136	S—S	208
H—I	161[a]	C—I	176	N—N	140	O—O	132		
H—N	98	C—N	147	N—O	136	F—F	141[a]		
H—O	94	C—O	143	O—H	94	Cl—Cl	199[a]		
H—S	132	C—S	181	O—C	143	Br—Br	228[a]		
Multiple covalent bonds[b]									
C═C	134	C≡C	120	C═O	121	O═O	121[a]	N≡N	113[a]

[a]Exact value.

[b]We'll discuss multiple covalent bonds in Section 7.5.

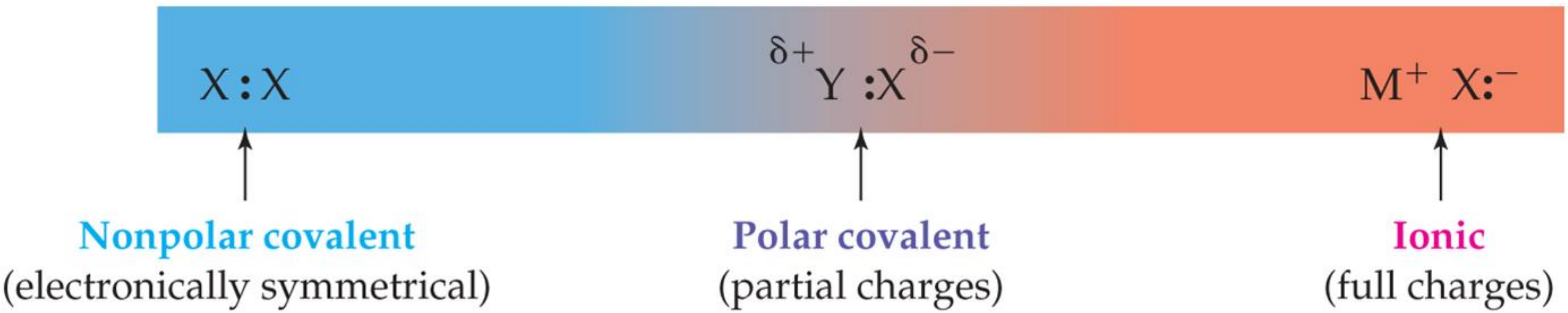
*Cleaving all bonds in multiple bond

Polar Covalent Bonds: Electronegativity

Electronegativity: atom's ability to attract shared electrons in a covalent bond

partial - charge $\delta -$

partial + charge $\delta +$



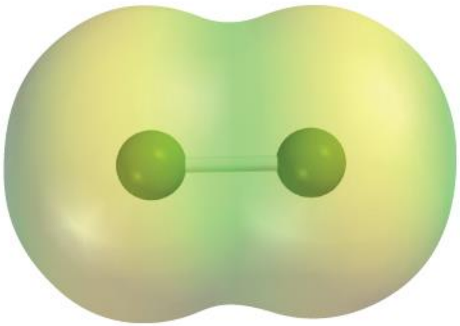
Polar covalent bonds have an unsymmetrical electron distribution in which the bonding electrons, shown as dots, are attracted more strongly by one atom than the other.

Polar Covalent Bonds: Electronegativity

The two bonding electrons, shown here as dots, are symmetrically distributed between the two Cl atoms.

Cl:Cl

A nonpolar covalent bond. Yellow-green represents a neutral atom.



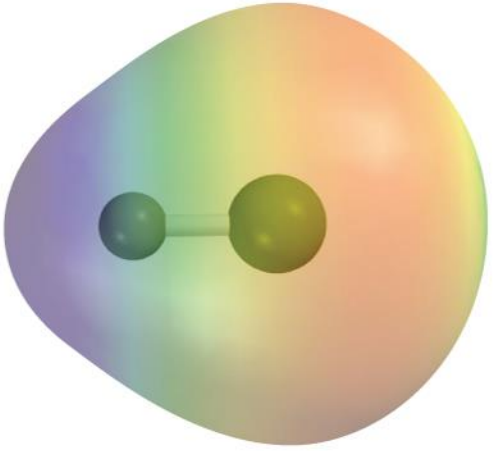
Electron distribution diagram –
yellow no charge, red - charge, blue + charge

Polar Covalent Bonds: Electronegativity

$^{\delta+}H-Cl^{\delta-}$

$[H\ {:}Cl]$

A polar covalent bond. The two bonding electrons (dots) are attracted more strongly by Cl than by H.

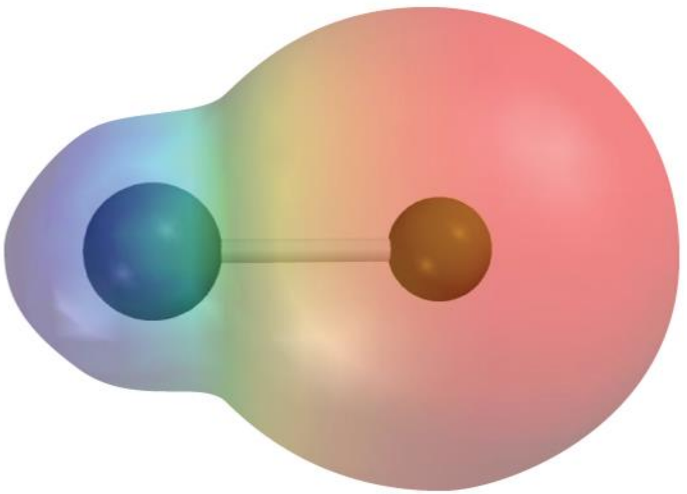


Electron distribution diagram –
yellow no charge, red - charge, blue + charge

Polar Covalent Bonds: Electronegativity

$Na^{+}Cl^{-}$

An ionic bond. Blue indicates a partial positive charge; red indicates a partial negative charge.

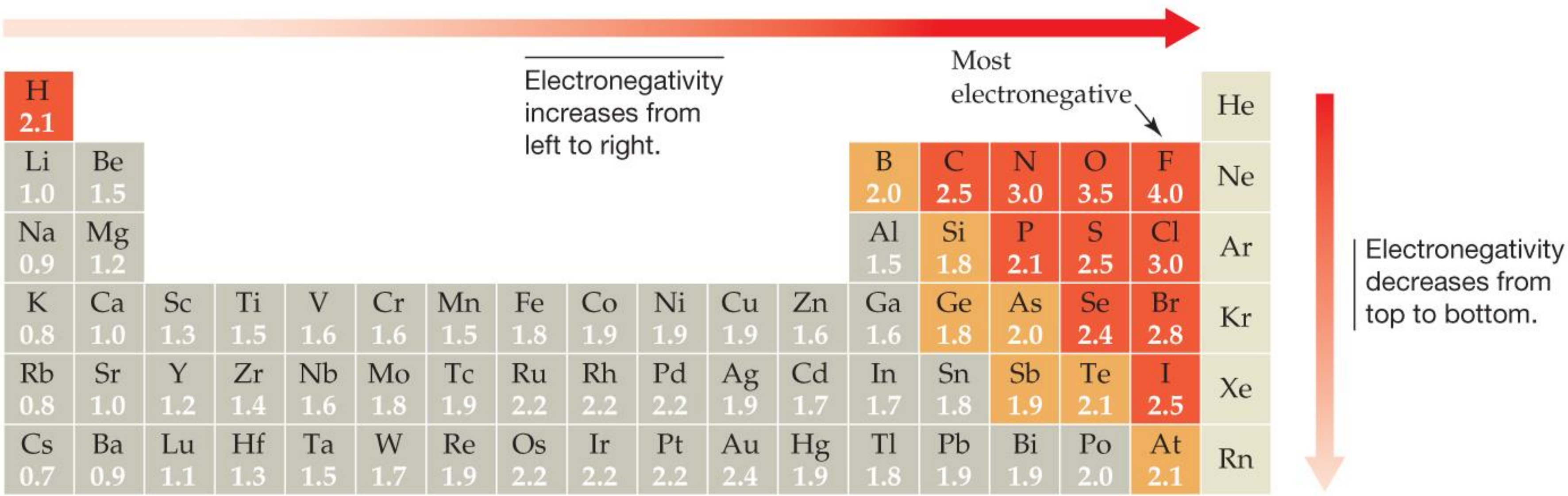


Electron distribution diagram –
yellow no charge, red - charge, blue + charge

Polar Covalent Bonds: Electronegativity

F is most electronegative (memorize this)

(H almost same EN as C)



HW #1: Electronegativity (EN, my abbreviation from another text)

F is most electronegative (memorize this)

(H almost same EN as C)

HW #2: Electronegativity: nonpolar, polar covalent, ionic ?

F is most electronegative (memorize this)
(H almost same EN as C)

HW #2: Electronegativity (EN, my abbreviation from another text)

F is most electronegative (memorize this)
(H almost same EN as C)

A Comparison of Ionic and Covalent Bonds

TABLE 7.3 Some Physical Properties of NaCl and HCl

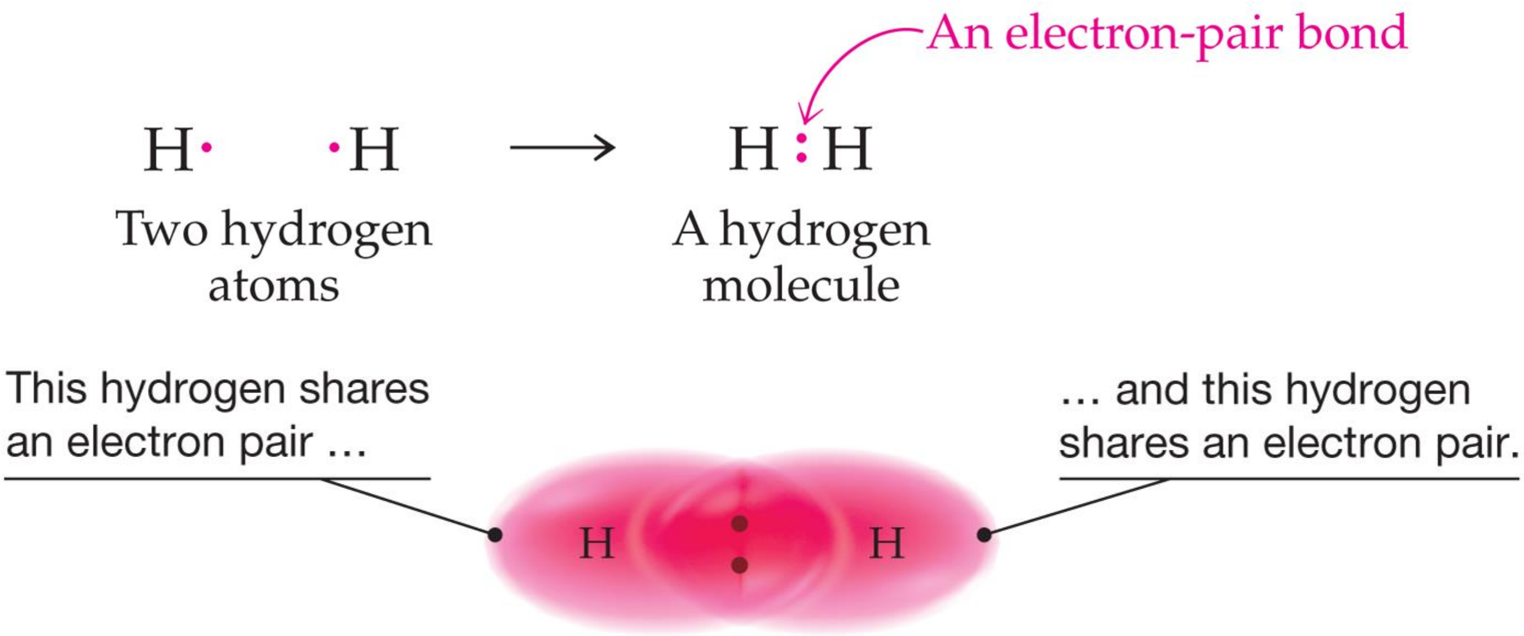
Property	NaCl	HCl
Formula mass	58.44 amu	36.46 amu
Physical appearance	White solid	Colorless gas
Type of bond	Ionic	Covalent
Melting point	801 °C	−115 °C
Boiling point	1465 °C	−84.9 °C

<u>ionic</u>	vs	<u>covalent</u>
High MP,BP		Low MP,BP

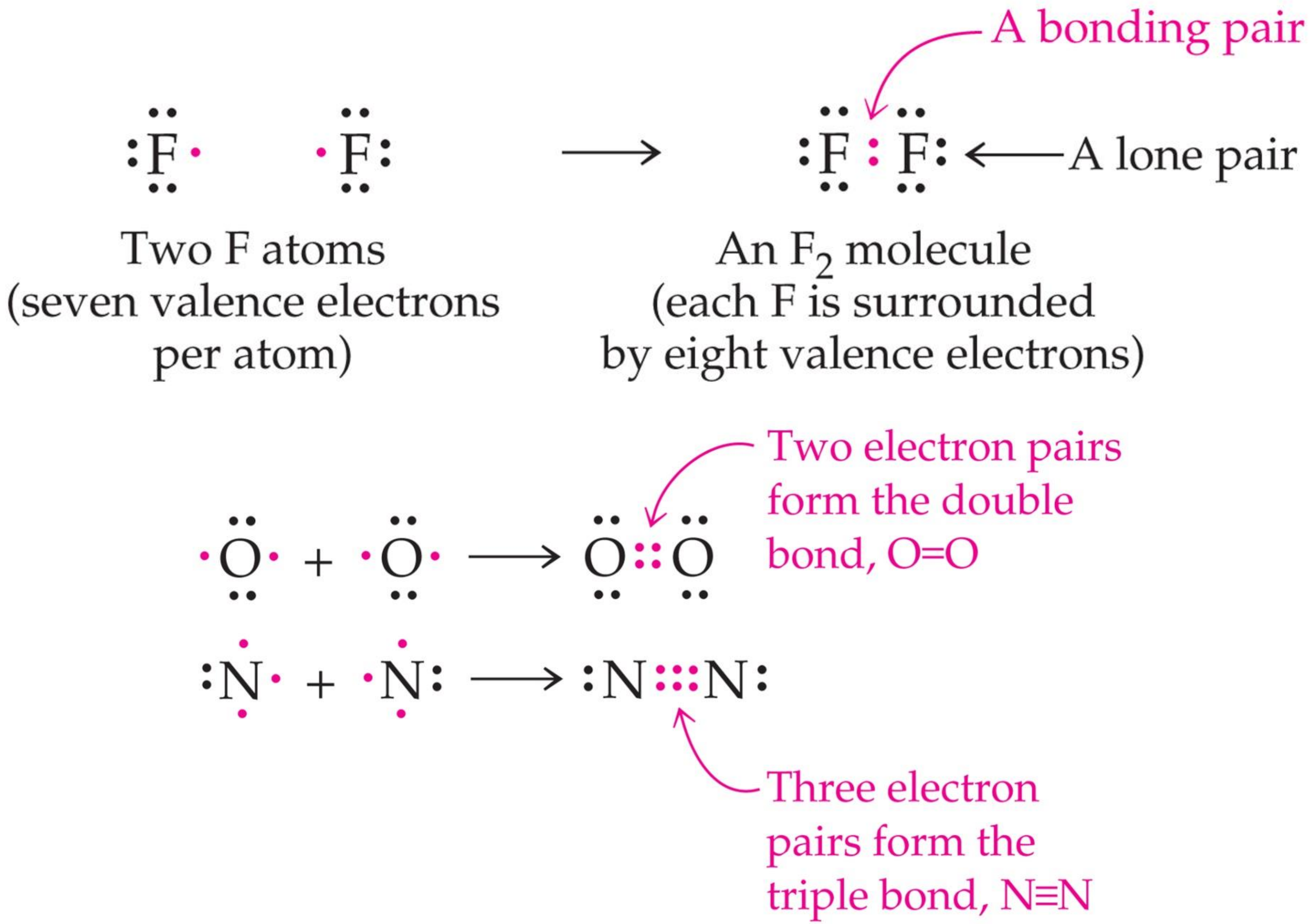
MP,BP have to do with interaction between molecules – <u>intermolecular forces (get in chapter 8)</u>

Electron-Dot Structures: The Octet Rule

Electron-Dot Structure (Lewis Dot Structure):
Represents an atom's valence electrons by dots

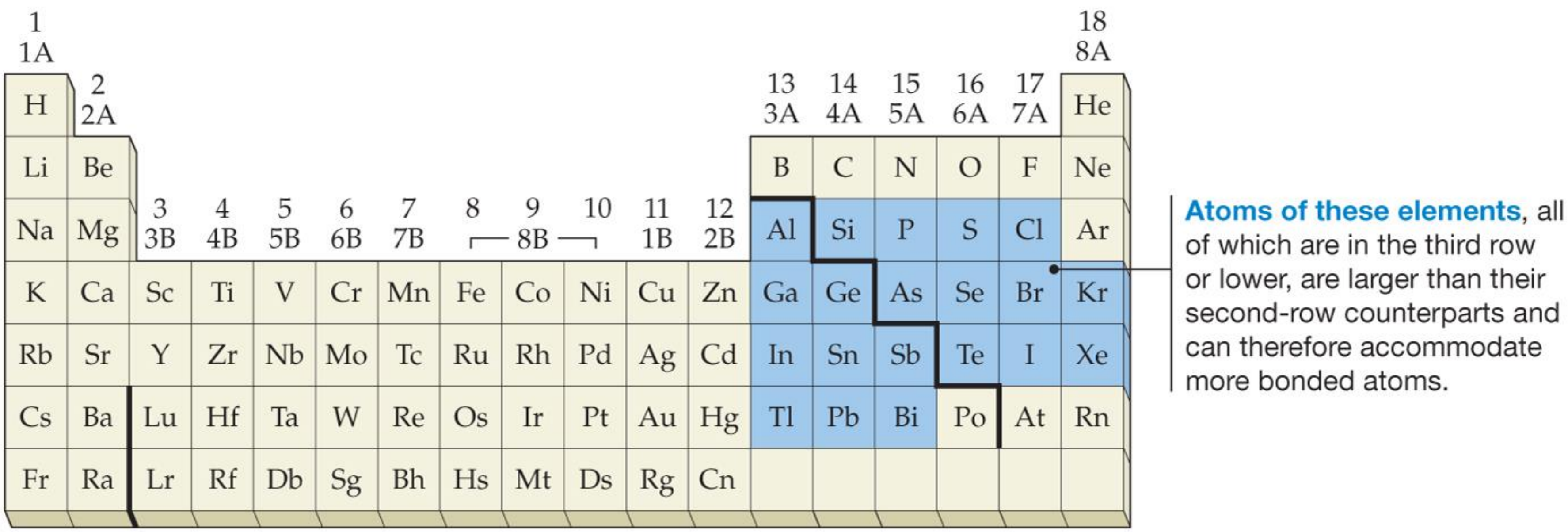


Electron-Dot Structures: The Octet Rule



Procedure for Drawing Electron-Dot Structures

The Octet Rule



3rd period (row) & higher main group elements have d subshell which allows for expansion of octet.

Procedure for Drawing Electron-Dot Structures (or Lewis Dot Structures)

Step 1: Valence Electrons

- Add up valence electrons for all atoms in the molecule. [ex: H_2O (2*1) + 6 = 8 e]
- Add one electron for each negative charge in an anion, or subtract one electron for each positive charge in a cation.

[ex: SO_4^{-2} 6 + (4*6) + 2 = 32e]

valence e for atom = group # for main group elements (1A to 8A-American system)

Procedure for Drawing Electron-Dot Structures

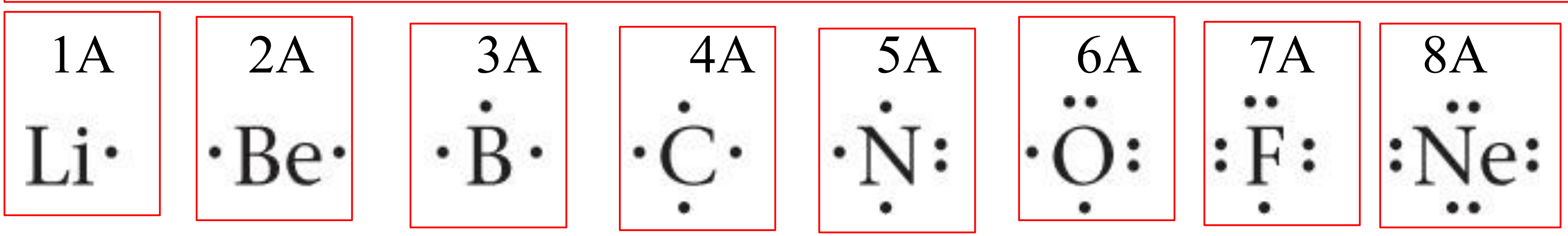
Step 2: Connect Atoms

- Draw lines between all atoms to represent bonds between atoms.
- Hydrogen and halogens usually form only one bond.
- Elements in third row and lower can expand octet (can have more than 8 electrons because have d subshell available even if d subshell is empty)

Lewis Structures of Atoms

- We use dots around the symbol to represent valence electrons. (4 walls – put one electron on each wall until run out of walls then double up electrons on walls with a dot already on it)

one electron wall forms one bond – other atom supplies other electron
two electron wall does not normally form a bond.



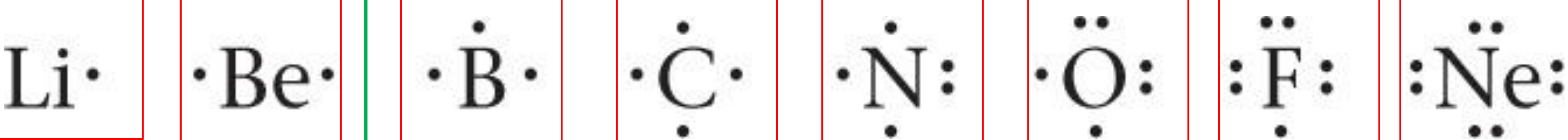
How many bonds for each of the above main group elements ?

Lewis Structures of Atoms

- We use dots around the symbol to represent valence electrons. (4 walls – put one electron on each wall until run out of walls then double up electrons)

How many covalent bonds for each of the main group elements ?

usually ionic bond		3 bond	4 bond	3 bond	2 bond	1 bond	no bond
1A	2A	3A	4A	5A	6A	7A	8A



Procedure for Drawing Electron-Dot Structures (below table from your text – same idea as last slide)

TABLE 7.4 Covalent Bonding for Second-Row Elements

Group	Number of Valence Electrons	Number of Bonds	Example
3A	3	3	BH_3
4A	4	4	CH_4
5A	5	3	NH_3
6A	6	2	H_2O
7A	7	1	HF
8A	8	0	Ne

End 11/8 Friday F section

End 11/11 Monday D section

Procedure for Drawing Electron-Dot Structures

Step 3: Put octets on all atoms

- Complete all atom's octet (bond single line counts as 2 electrons) (except for hydrogen – H, He only gets duet).

(to get octet: use lone pair electrons if not enough bonds)

Procedure for Drawing Electron-Dot Structures

Step 4: Check # electrons in your Lewis Dot Structure:

 # electrons in your structure = # valence electrons

done.

End 11/8/19 Friday G section

Procedure for Drawing Electron-Dot Structures

Step 5: Put in Multiple Bonds. OR Put in lone pairs.

If the number of electrons from step 4 does not match up, put in multiple bonds OR add lone pairs to central atom.

(each multiple bond decreases number of e in structure by 2 e) (each lone pair increase e in structure by 2e)

Go back to step 3 & redo iteratively until done. (done: # e in structure = # valence e)

Procedure for Drawing Electron-Dot Structures

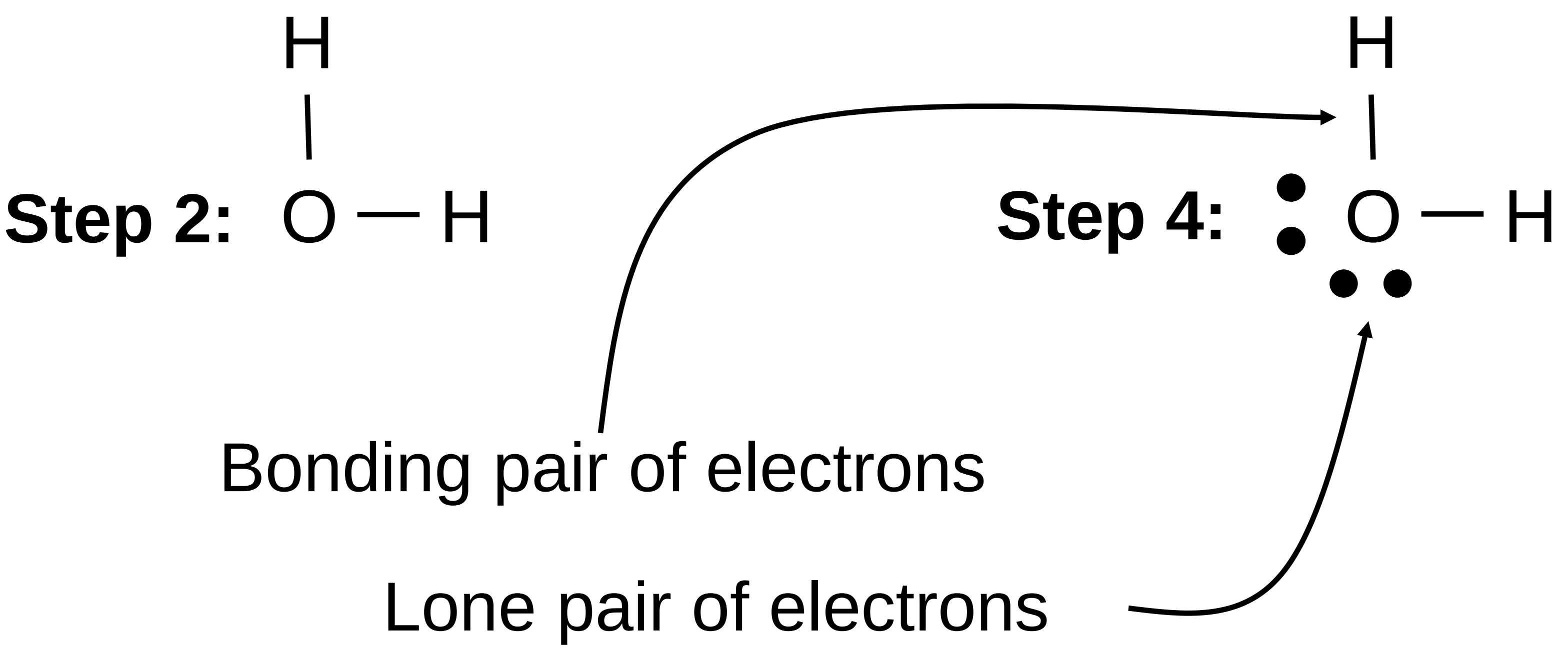
Draw an electron-dot structure for $\mathbf{H_2O}$.

Step 1: 2(1) + 6 = 8 valence electrons

Procedure for Drawing Electron-Dot Structures

Draw an electron-dot structure for **H_2O**.

Step 1: 2(1) + 6 = 8 valence electrons



Procedure for Drawing Electron-Dot Structures

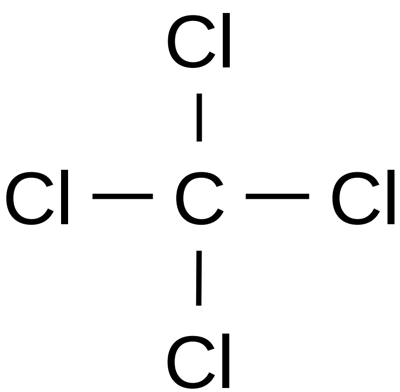
Draw an electron-dot structure for $\mathbf{CCl_4}$.

Procedure for Drawing Electron-Dot Structures

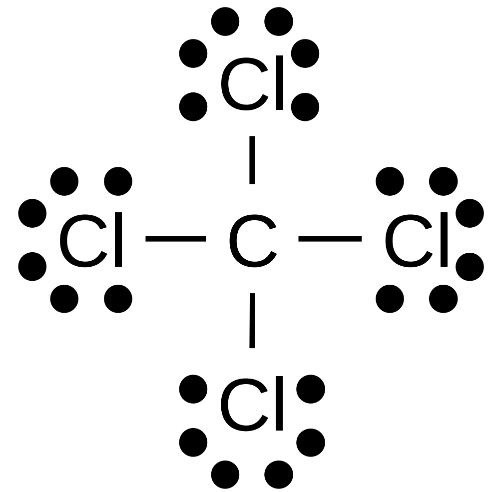
Draw an electron-dot structure for $\mathbf{CCl_4}$.

Step 1: 4 + 4(7) = 32 valence electrons

Step 2:



Step 3:



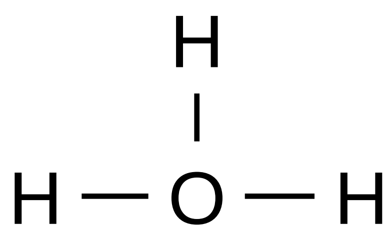
Procedure for Drawing Electron-Dot Structures

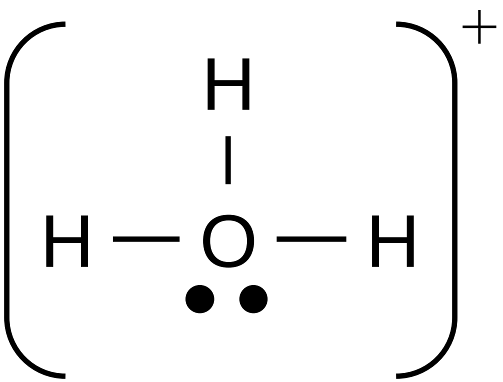
Draw an electron-dot structure for $\mathbf{H_3O^{1+}}$.

Procedure for Drawing Electron-Dot Structures

Draw an electron-dot structure for $\mathbf{H_3O^{1+}}$.

Step 1: 3(1) + 6 – 1 = 8 valence electrons

Step 2: 

Step 4: 

End 11/13 Wed
section D

Procedure for Drawing Electron-Dot Structures

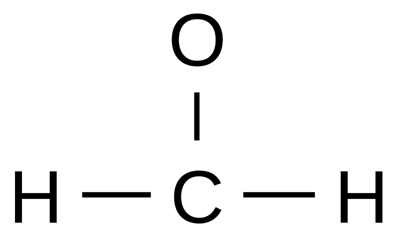
Draw an electron-dot structure for $\mathbf{CH_2O}$.

Procedure for Drawing Electron-Dot Structures

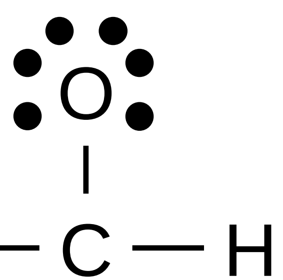
Draw an electron-dot structure for **CH_2O**.

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

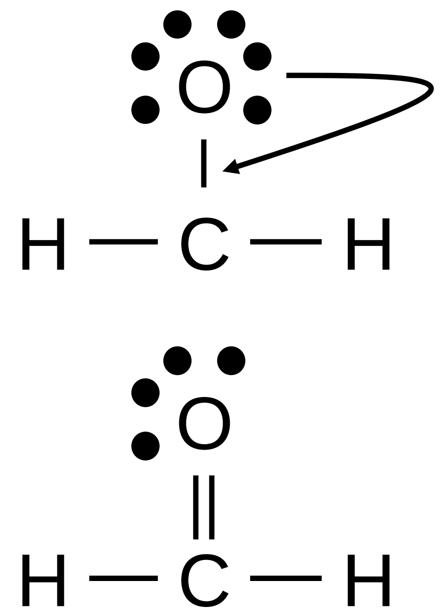
Step 2:



Step 3:



Step 5:



Procedure for Drawing Electron-Dot Structures

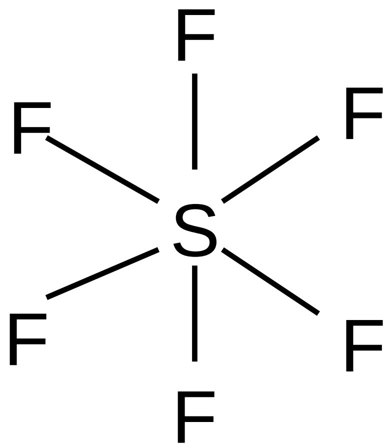
Draw an electron-dot structure for **SF_6**.

Procedure for Drawing Electron-Dot Structures

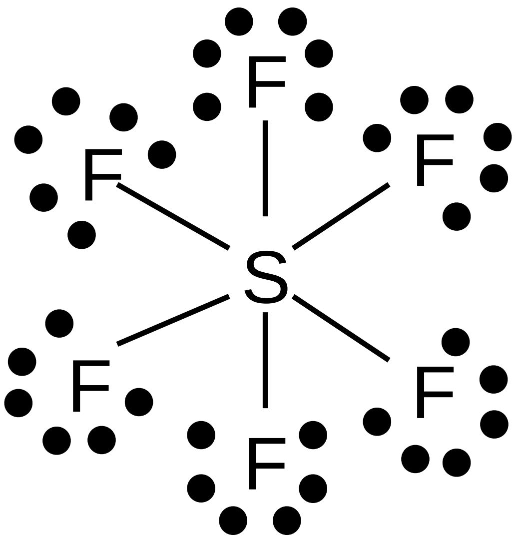
Draw an electron-dot structure for **SF_6**.

Step 1: 6 + 6(7) = 48 valence electrons

Step 2:



Step 3:



Procedure for Drawing Electron-Dot Structures

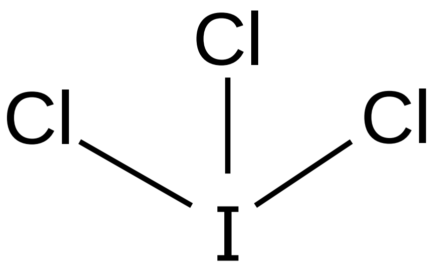
Draw an electron-dot structure for $\mathbf{ICl_3}$.

Procedure for Drawing Electron-Dot Structures

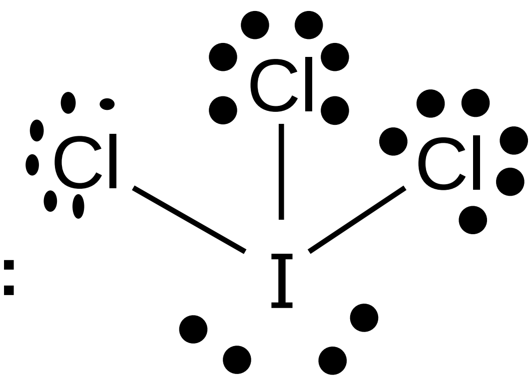
Draw an electron-dot structure for **ICl_3**.

Step 1: 7 + 3(7) = 28 valence electrons

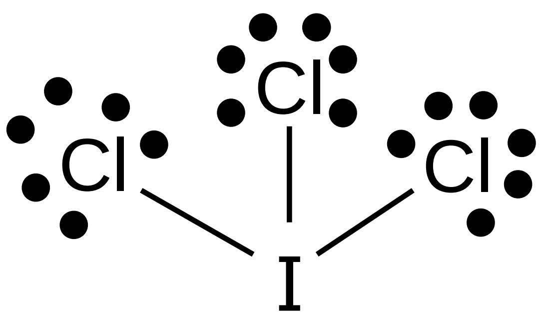
Step 2:



Step 4:



Step 3:



HW #3: Draw an electron-dot (Lewis Dot) structure

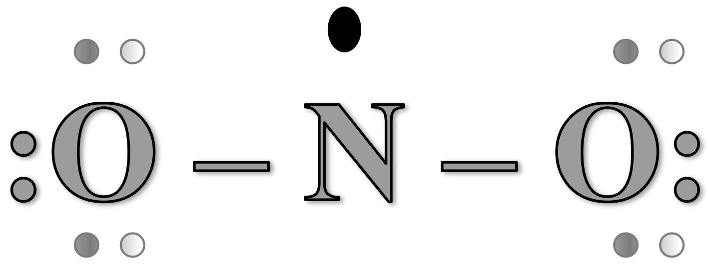
End 11/11 Monday F section

Drawing Electron-Dot Structures for Radicals (on list of topic left off from common syllabus)

- Lewis Dot Structure with unpaired electron. These are called *radicals*, or *free radicals*.
- Drawing electron-dot structures for radicals follows the steps we have seen.
- There will always be an unfilled octet on one atom.

Drawing Electron-Dot Structures for Radicals (not responsible)

- As an example, NO_2
  - Each oxygen provides 6 electrons.
  - The nitrogen provides 5 electrons.
  - Thus, there are 17 electrons.
  - Note the lone electron on the nitrogen.



End 11/11 Monday G section

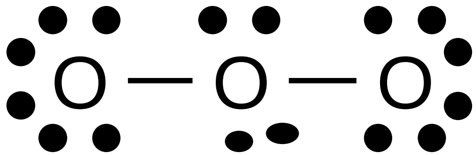
Electron-Dot Structures and Resonance

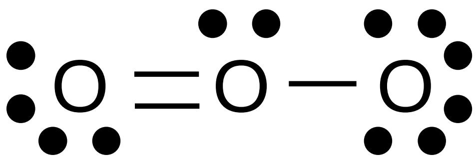
Draw an electron-dot structure for $\mathbf{O_3}$.

Step 1: 3(6) = 18 valence electrons

Step 2: O — O — O

10x2= 20 e – too many e

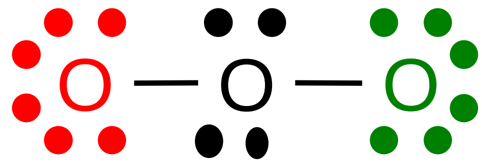
Step 3 

Step 5 **:** 

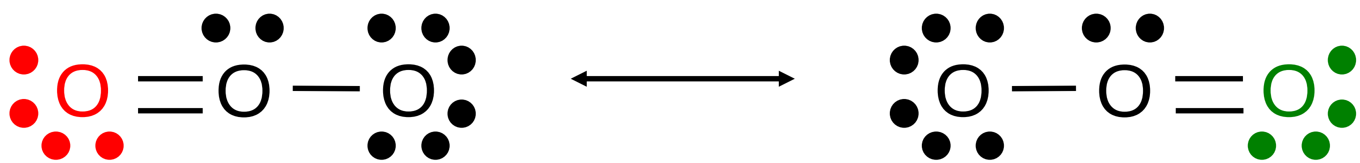
Electron-Dot Structures and Resonance

Move a lone pair from this oxygen?

Step 4:



Or move a lone pair from this oxygen?



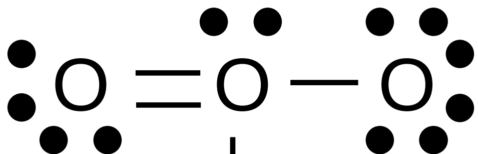
Resonance

End section F
11/13

Formal Charges

$$\textbf{Formal charge} = \left(\begin{array}{c}\text{\# of}\\ \text{valence } e^{-}\\ \text{in free atom}\end{array}\right) - \frac{1}{2}\left(\begin{array}{c}\text{\# of}\\ \text{bonding}\\ e^{-}\end{array}\right) - \left(\begin{array}{c}\text{\# of}\\ \text{nonbonding}\\ e^{-}\end{array}\right)$$

Calculate the formal charge on each atom in O_3.



$$6 - \frac{1}{2}(4) - 4 = 0 \qquad 6 - \frac{1}{2}(6) - 2 = +1 \qquad 6 - \frac{1}{2}(2) - 6 = -1$$

HW #4: Draw an electron-dot (Lewis Dot) structure. Give formal charge. Show all resonance structures.

$$\textbf{Formal charge} = \left(\begin{array}{c} \text{\# of} \\ \text{valence } e^- \\ \text{in free atom} \end{array} \right) - \frac{1}{2} \left(\begin{array}{c} \text{\# of} \\ \text{bonding} \\ e^- \end{array} \right) - \left(\begin{array}{c} \text{\# of} \\ \text{nonbonding} \\ e^- \end{array} \right)$$

End 11/14 Thursday section D

End 11/13 Wednesday section G