Structure & Properties:

Chemical Bonding

Chemistry 12 (4.1)

Ionic and Covalent Bonds

Metallic Bonding

Lewis Structures

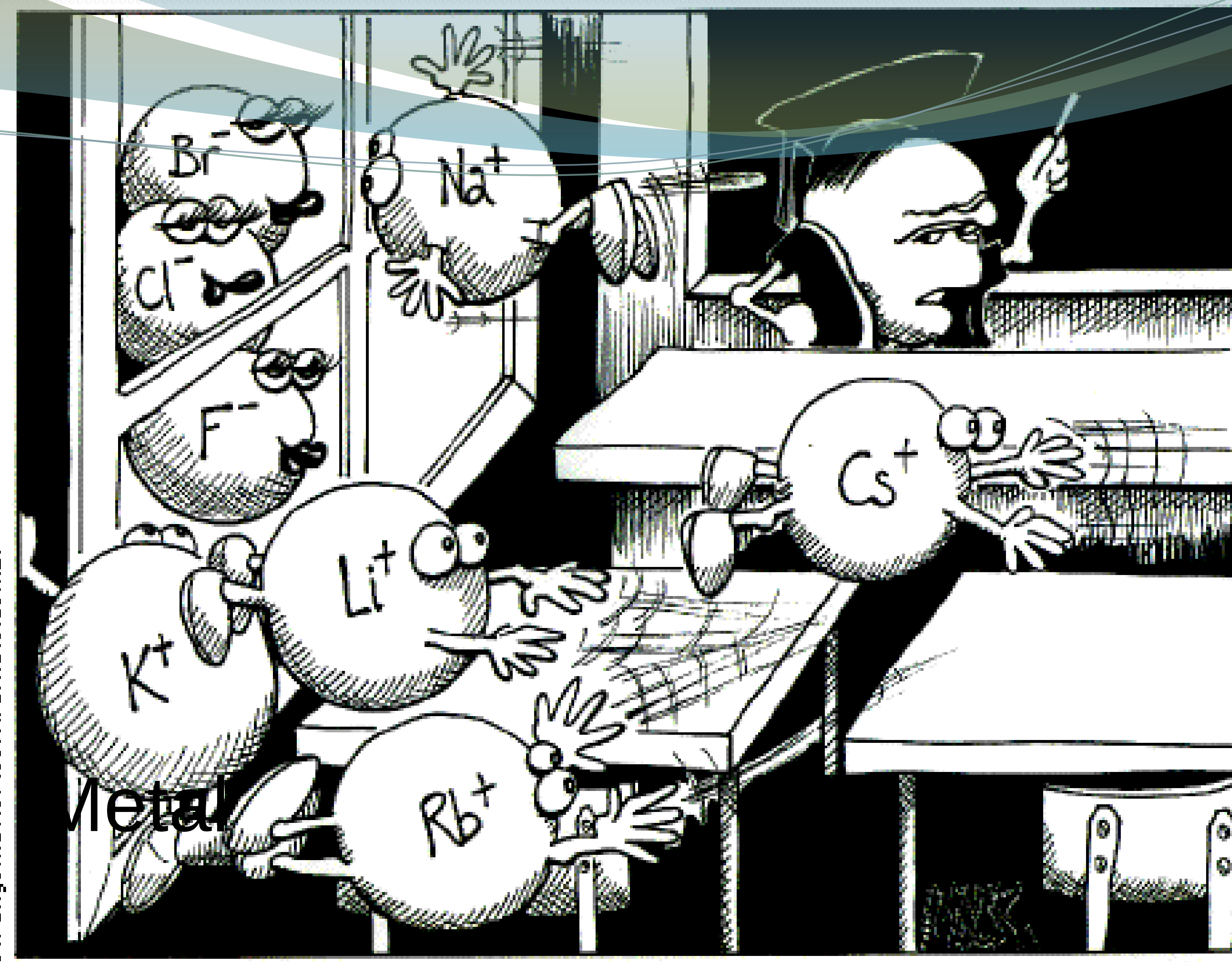
Chemical Bonds

Bonding

- Bonds are formed between atoms by the transfer or sharing of electrons
- Atoms acquire stable octets by forming bonds (8 electrons)
- Three types of bonding:
  - Ionic - metal & non-metal
  - Covalent - non-metal & non-metal
  - Metallic - metals only

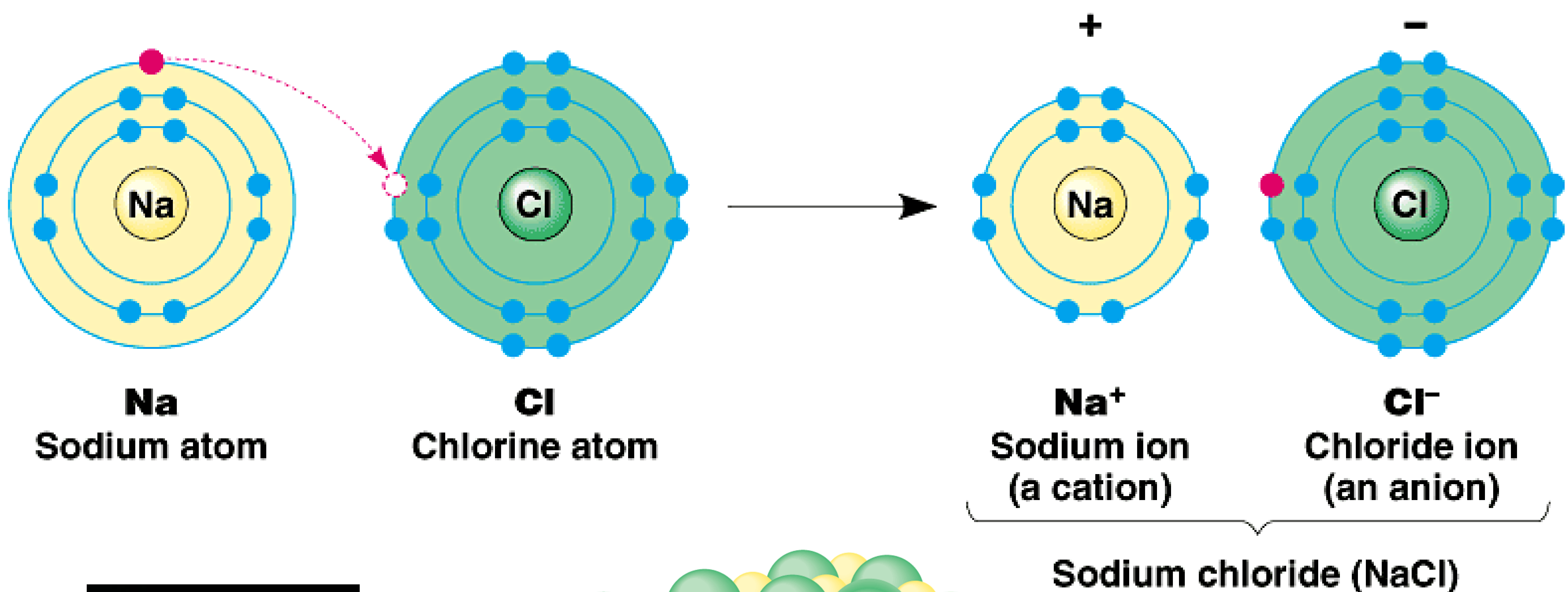
Ionic Bonding

- Occurs between a Metal and Non-metal
- Transfer of electrons
- Ions are formed in the process
- “Bond” is an electrostatic attraction between ions
- Difference in electronegativity more than 1.7

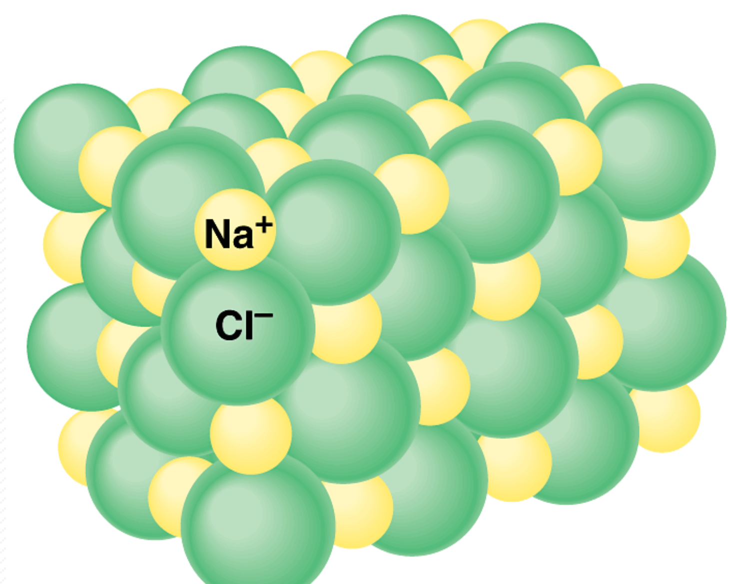


“Perhaps one of you gentlemen would mind telling me just what it is outside the window that you find so attractive...?”

Ionic Bonds – occur when one atom donates or gives up one or more electrons



Salt crystals

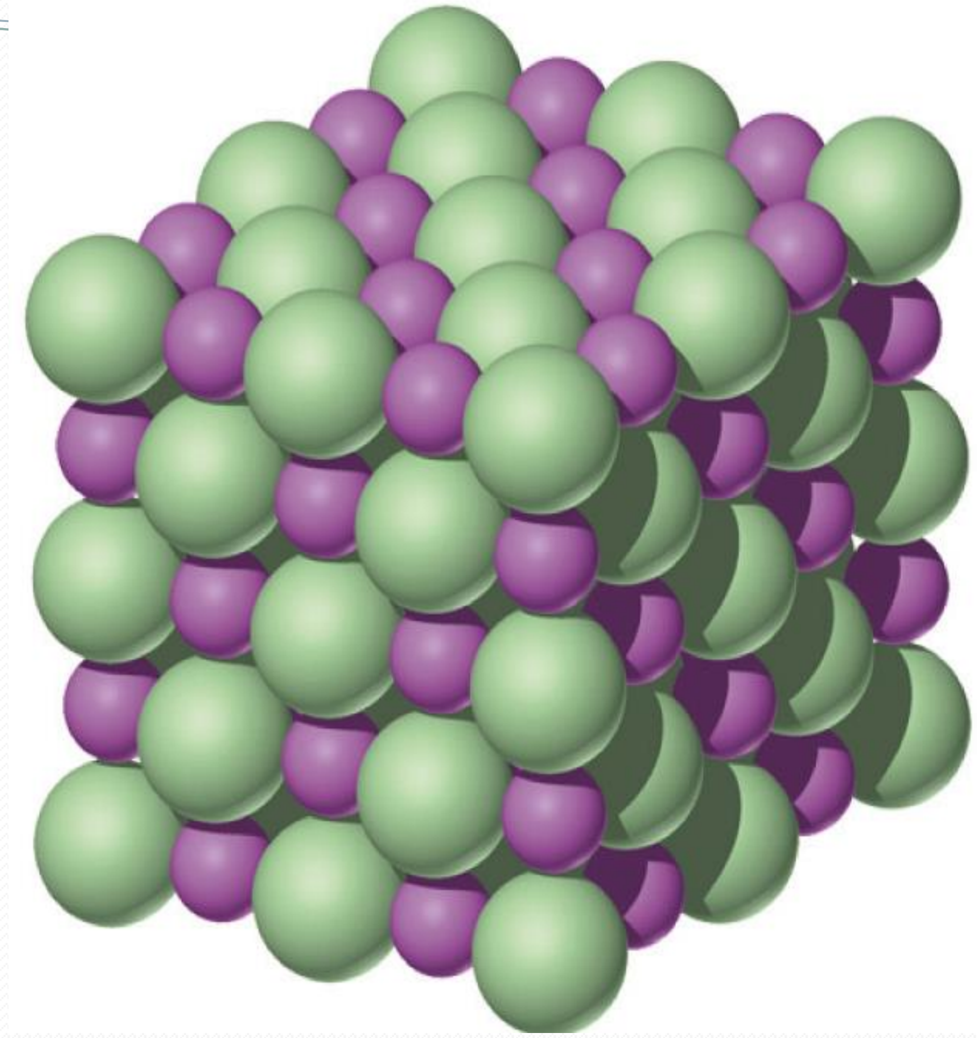


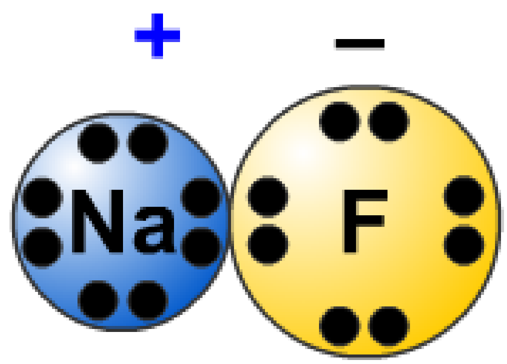
Ionic Compound (Na^+Cl^-)

Opposite charges attract to form ionic bonds

Ionic Bonding

- Ionic bonds such as NaCl do not consist of just one Na^+ ion bonded to one Cl^- ion; rather, ionic bonds represent the relative ratios of these ions in a huge lattice of thousands or millions of ions (i.e. – if there are 5,000 Na^+ ions, there are 5,000 Cl^- ions)

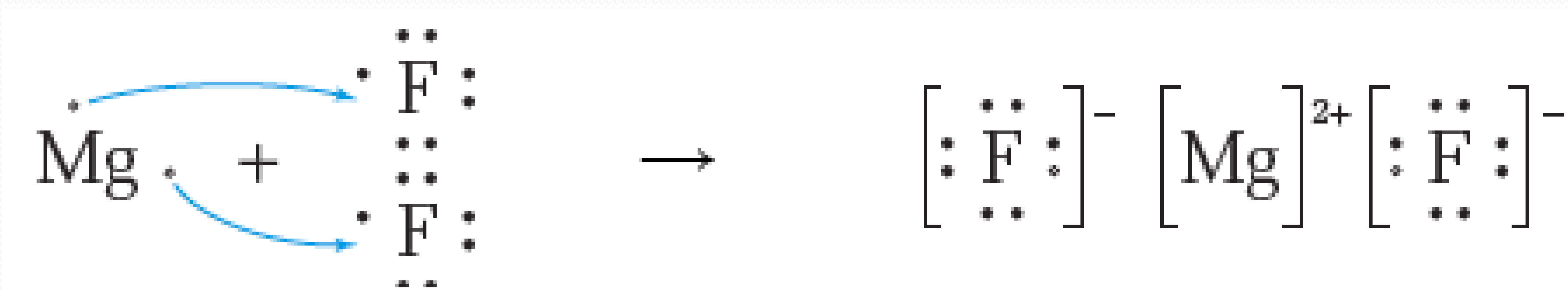




Properties of Ionic Compounds

- Properties of IONIC compounds -
- mostly solids at room temperature
- do not conduct electricity in solid state
- easily dissolve in water to form electrolytes, so can conduct electricity in aqueous and liquid state
- hard and brittle
- high melting and boiling points

cation	anion	compound
Ca^{+2}	Cl^{-1}	$CaCl_2$
Ba^{+2}	O^{-2}	BaO
K^{+1}	S^{-2}	K_2S
Fe^{+3}	Br^{-1}	$FeBr_3$
Cr^{+3}	O^{-2}	Cr_2O_3



Metallic Bonding

- Scientists developed a model for how metals bonds based on the physical properties:
  - conductivity requires charged particles
  - hardness implies strong bonding
  - malleability implies regularity in the structure
  - ductility implies regularity in the structure

Delocalized Electrons

- Electrons are not confined to a particular location and are free to move around the structure
- The negative electrons are attracted to the positive protons, keeping the lattice structure together

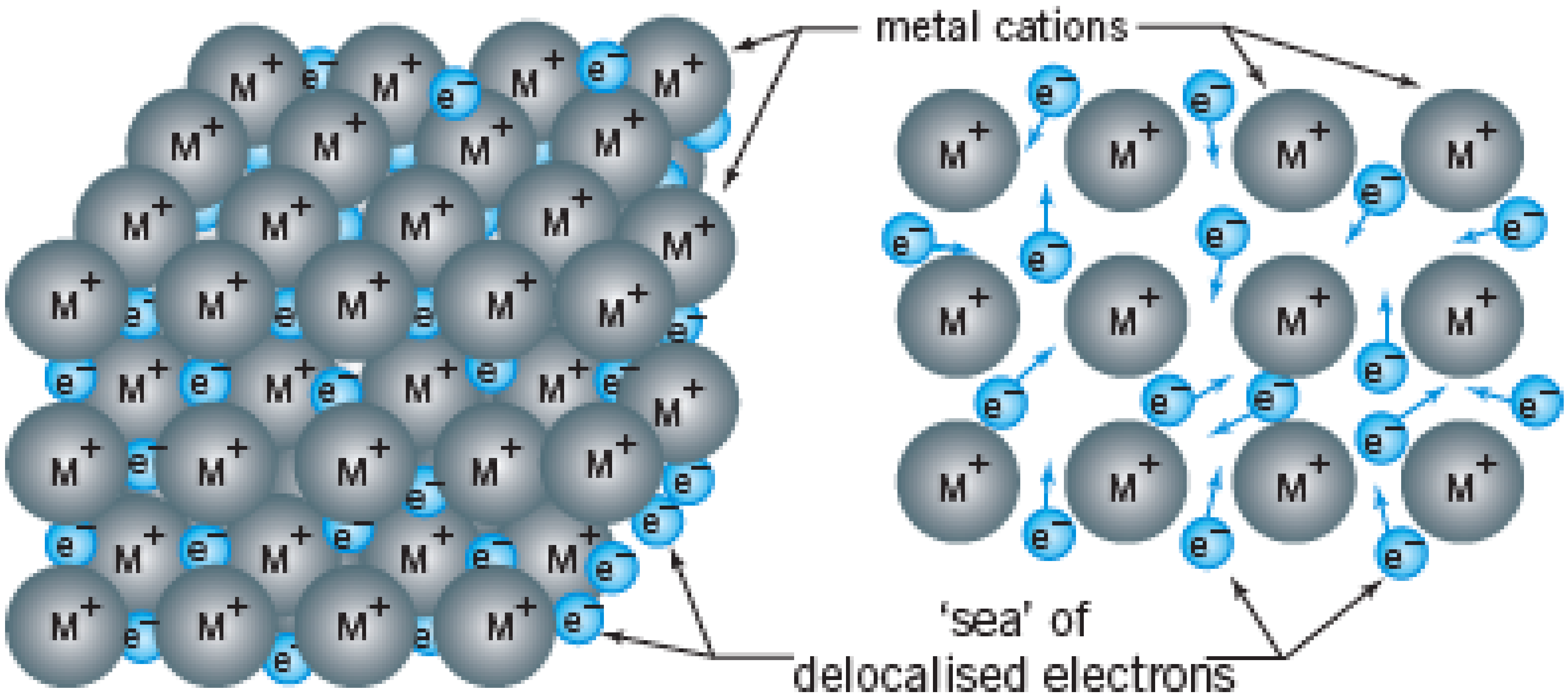
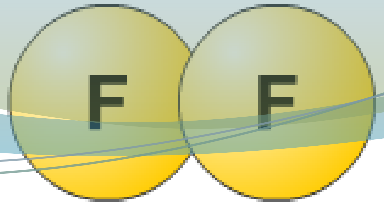
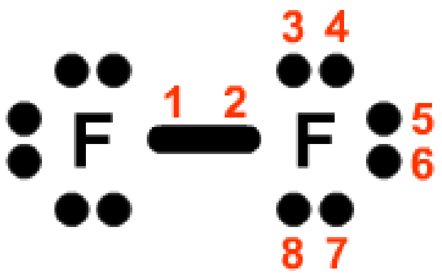


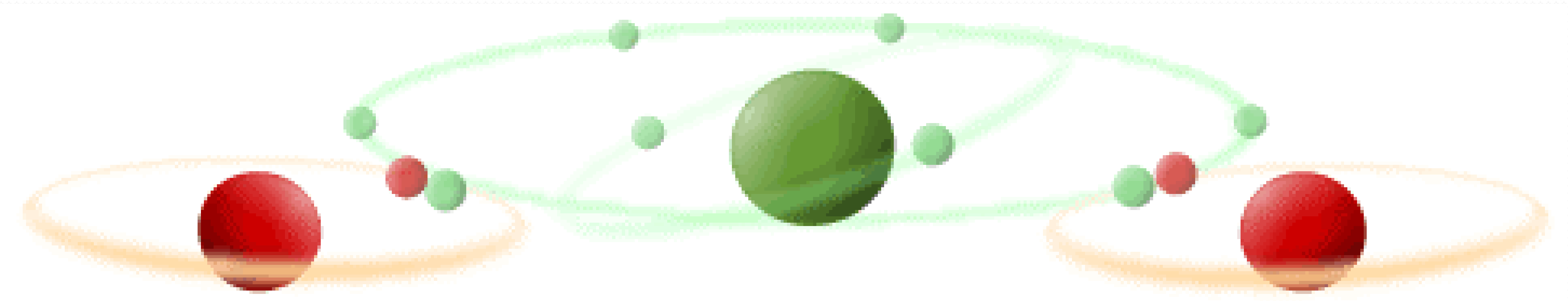
Figure 2.2.1 The structure of a metal may be described as a lattice of cations surrounded by a 'sea' of electrons.

Covalent Bonding



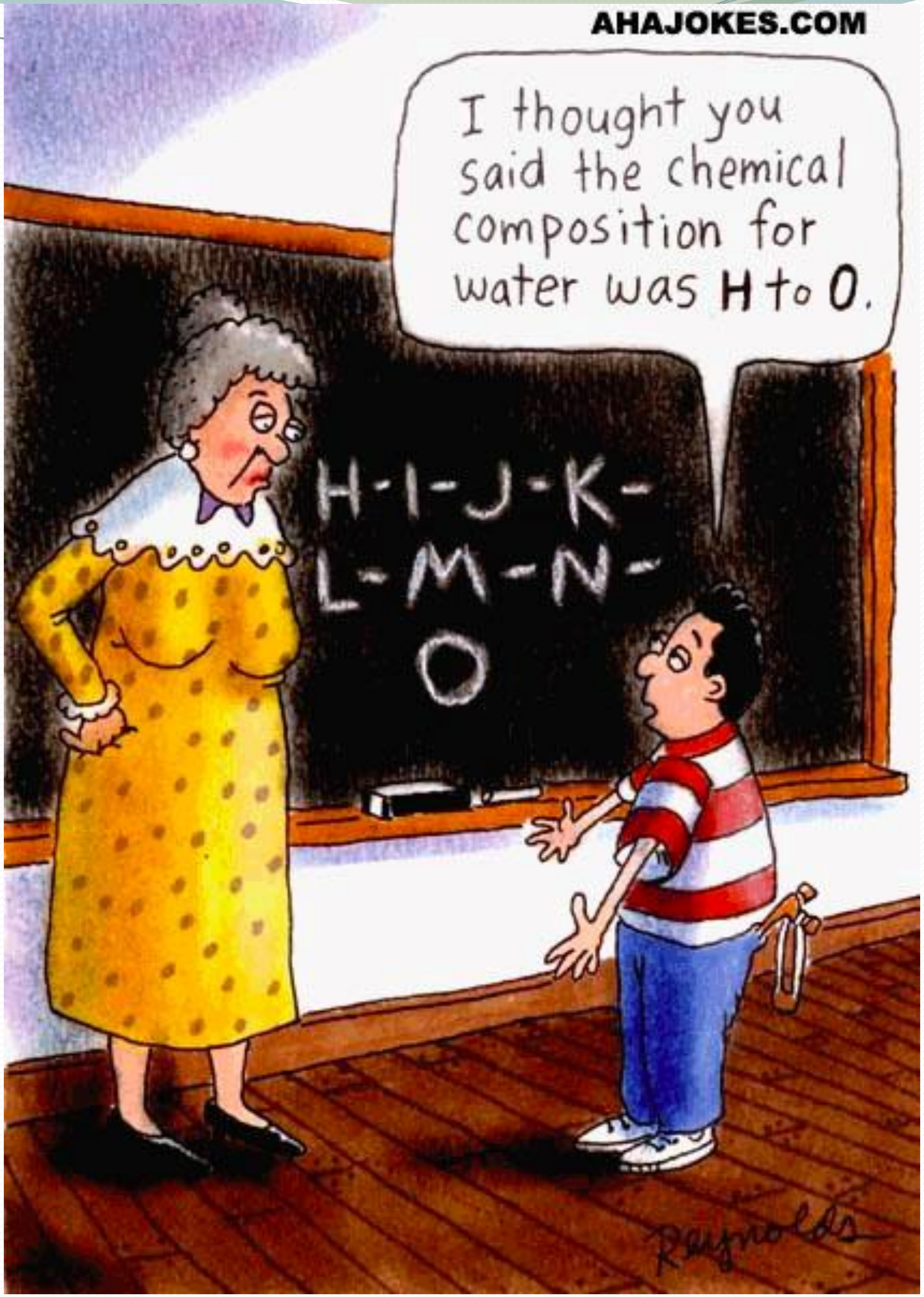


- Atoms with half filled atomic orbitals overlap and share the same space, forming a new orbital
- The formation of a new orbital has lower energy than the original orbitals
- Consider the covalent bonding of hydrogen and oxygen in water, for example:
  - H has 1 valence electron, while O has 6 valence electrons

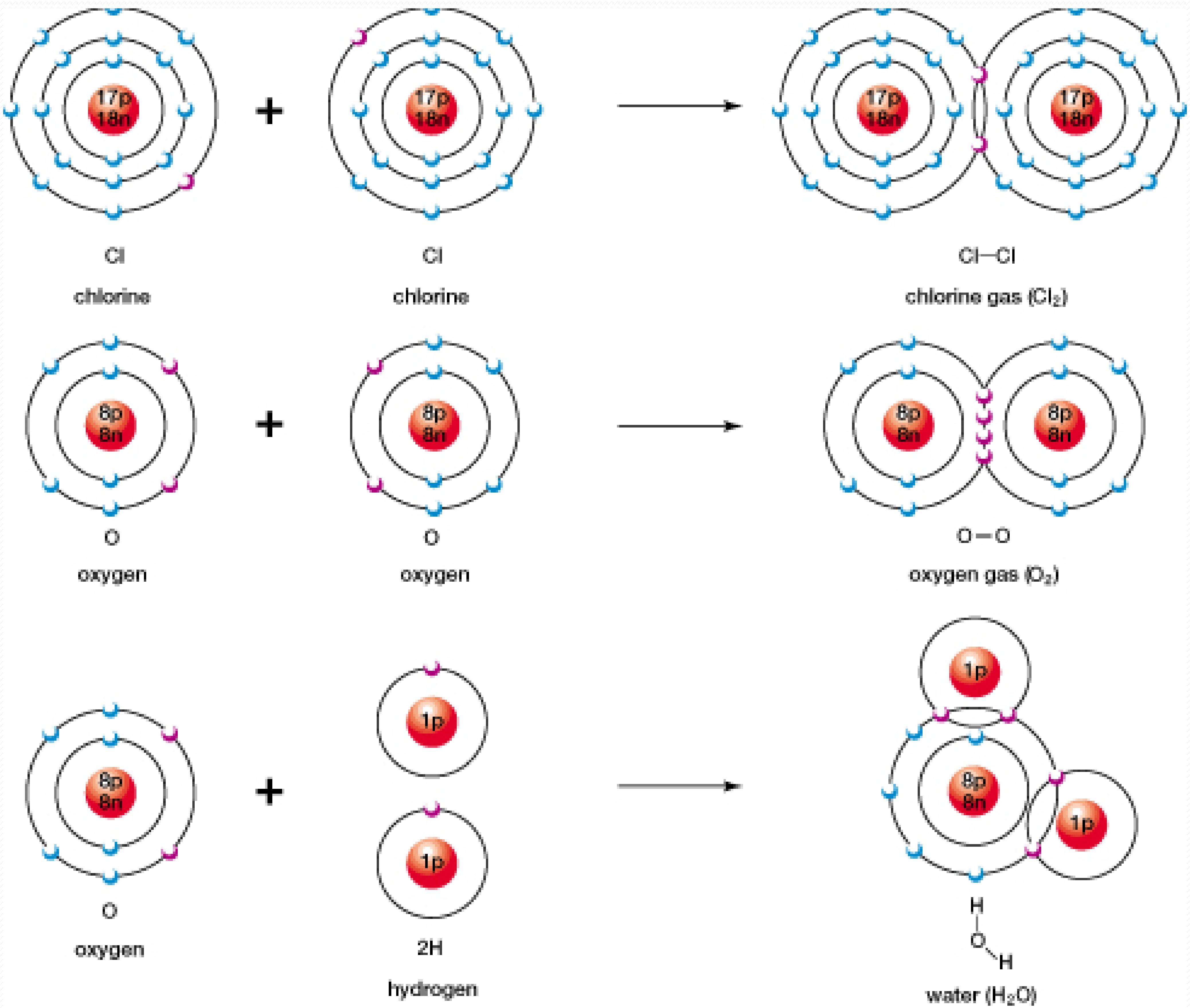


Covalent

- Non-metal and non-metal
- Sharing of electrons
- No ions formed in the process
- Bond is shared electrons between the two atoms (2, 4 or 6 electrons)
- Difference in electronegativity less than 1.7



Covalent Bonds – involve a sharing of a pair of valence electrons between atoms.



Covalent Properties

- solid, liquid or gas at room temperature
- do not conduct electricity in solid or liquid states
- may or may not dissolve in water,
  but cannot conduct electricity in aqueous state
- soft, waxy or flexible
- stable at high temperatures; do not easily decompose upon heating, but will react in chemical reactions

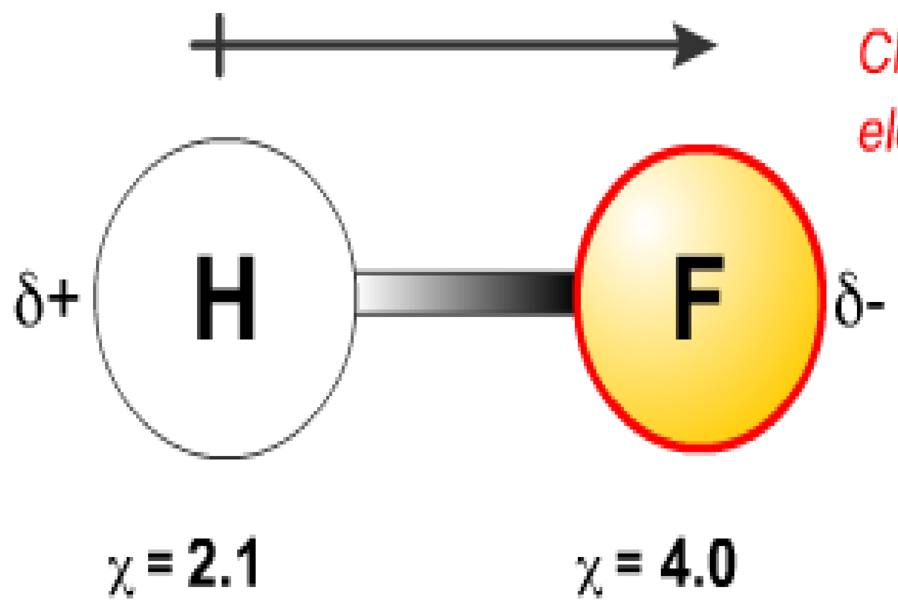
Electronegativity

- A value that describes the ability for a element to attract electrons. The higher the electronegativity, the more attraction they have for electrons.

Fluorine is more electronegative than hydrogen. Therefore, the electrons in the bond are attracted towards the fluorine atom.

Click the atom to change the element

The hydrogen atom has a partial positive charge.

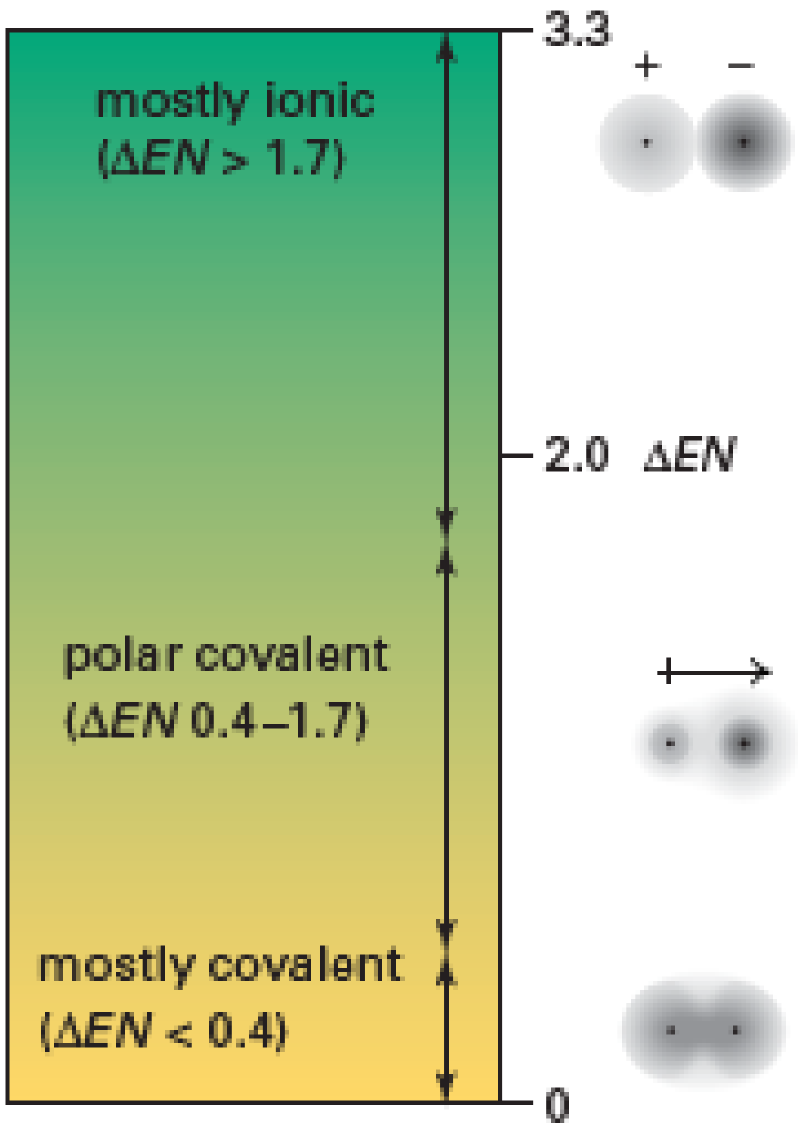


The fluorine atom has a partial negative charge.

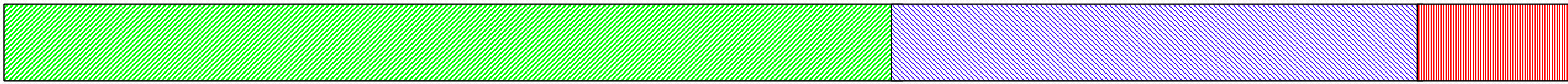
This is called a **polar bond** and produces a **dipole moment**, shown by the arrow above the bond.

Bond Polarity

- In pure covalent bonds (found in diatomic molecules), electrons are shared equally
- Anything other than diatomic molecules do not share electrons equally, because the electronegativity of the atoms connected to the bond are different
- If the difference is greater than 1.7, the bond is considered an ionic bond
- If the difference is between zero and 1.7, the bond is considered a polar covalent bond



Ionic vs. Covalent



- <u>between 0.0 – 0.4</u> **non-polar covalent bond** – bonding electrons are shared equally

(H_2, O_2, F_2, Br_2, I_2, N_2, Cl_2) – they form a “7” on the table.

- <u>between 0.4 – 1.7</u> **polar covalent bond** – the more electronegative atom attracts the shared electron pair more strongly than the less electronegative atom (ex: HCl)

- <u>between 1.7 – 4.0</u> **ionic bond** – one atom completely loses its valence electrons and the other (more electronegative) atom gains them (ex: NaCl)

Bond Length and Strength

- In general, the length of bonds decrease as the number of bonds increase
- In general the strength of bonds increase as the number of bonds increase

TABLE 2.3.1 BOND DISSOCIATION ENTHALPIES AND BOND LENGTHS OF SOME BONDS

Bond	Bond dissociation enthalpy (kJ mol^{-1})	Bond length (pm, 1 pm = 10^{-12} m)
H–H	436	74
Cl–Cl	242	199
O=O	498	121
N≡N	945	109
C–C	364	154
C=C	602	134
C≡C	835	120
C–O	358	143
C=O	799	120

Homework!

- Please re-read Section 4.1 (pp.163-158) and answer:
- p.165-6 #1- 4
- p.169-170 #5-8
- p.171-2 #1-7

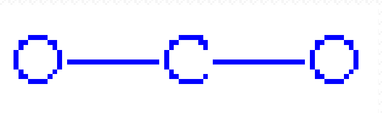
Steps for Drawing Lewis Structures

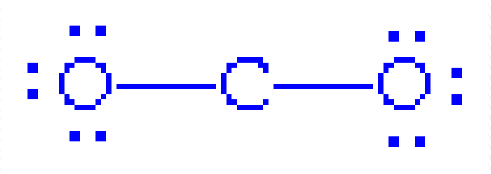
- Covalent Compounds

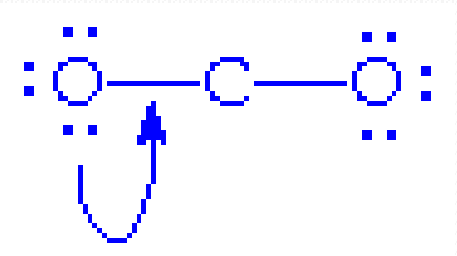
1. Count the total number of valence electrons in the molecule.
2. Place the atoms with the least electronegative atom as a central atom (if this applies). Most of the time, molecules will take the most symmetric shape possible.
3. Place single bonds between adjacent atoms. These are called bond pairs.
4. Electrons left = total electrons – bonds X 2
5. Place remaining electrons in pairs around atoms, starting with the terminal atoms. These are called lone pairs.
6. Check to see if all atoms have 8 electrons (exceptions: 2 for H, 4 for Be, 6 for B)
7. Move lone pairs to make bonds in order to satisfy the octet rule for all atoms.

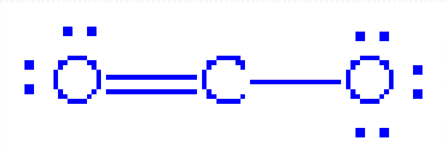
Example

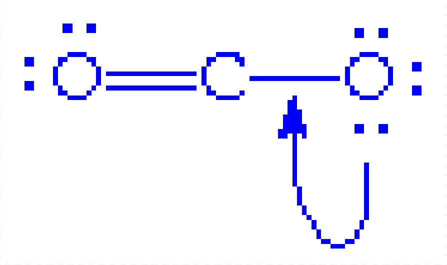
1. Carbon dioxide
2. C has 4 valence e^-, while O has 6 valence electrons
3. Total electron count = $16e^-$
4. Single bonds
5. 4 e^- used, 12 e^- left to place on terminal atoms
6. 16 e^- used. Carbon does not have octet
7. Move electron pairs to satisfy carbon, forming double bonds

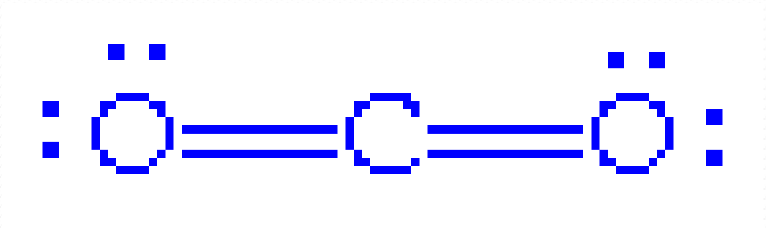












Try it!

- Cl_2 HCl

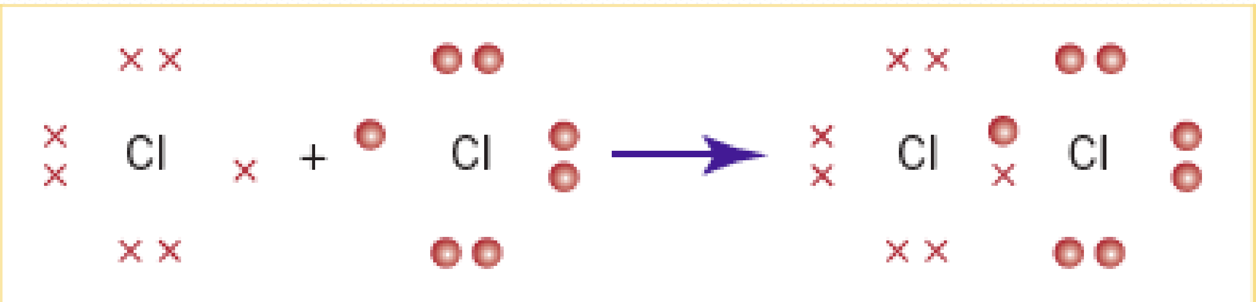


Figure 2.3.5 The Lewis (electron dot) structure of a chlorine molecule (Cl_2).

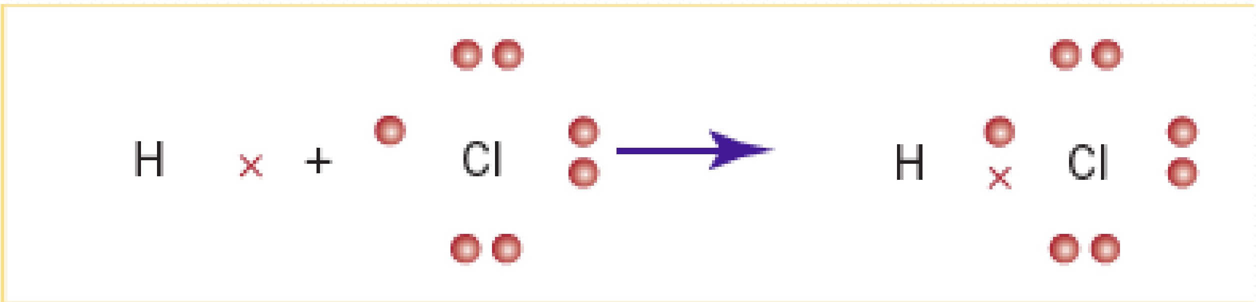
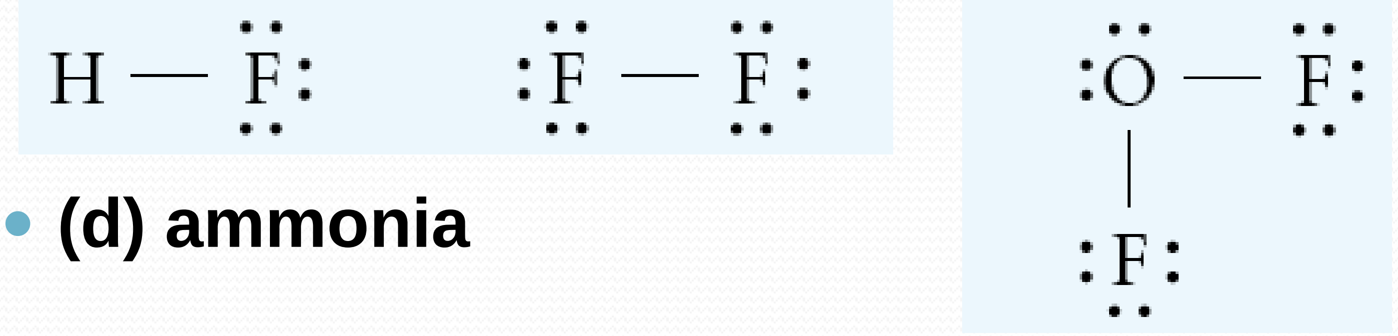


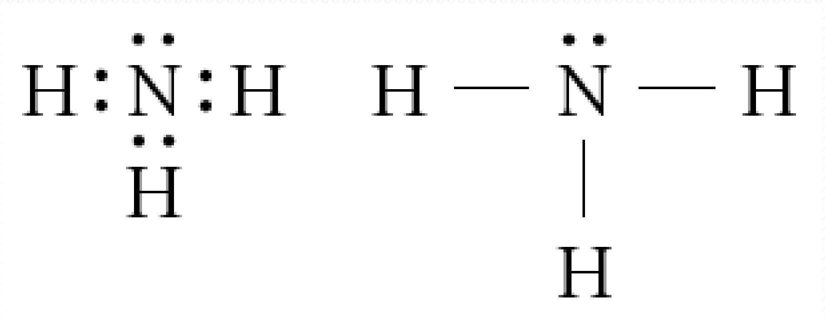
Figure 2.3.6 The Lewis structure of a molecule of hydrogen chloride (HCl).

Try some more!

(a) $HF_{(g)}$ **(b) $F_{2(g)}$** **(c) $OF_{2(g)}$**



- **(d) ammonia**



YOU MUST COUNT ELECTRONS WHEN DOING LEWIS STRUCTURES!!!

They are NOT pretty pictures...they are representing the chemical structures!!!

HOW TO DRAW A LEWIS STRUCTURE FOR A MOLECULE

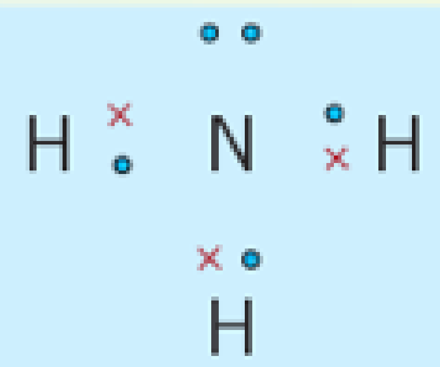
1. Determine how many valence electrons there are in each atom of the molecule.
   - e.g. NH_3
   - Valence electrons:
     Nitrogen 5
     Hydrogen 1

2. Determine how many more electrons are needed to fill the valence shell.
   - This is the number of bonds the atom will form.
   - Nitrogen: 3 bonds (central atom)
     Hydrogen: 1 bond each

3. Draw electron dot diagrams, pairing up all electrons except the number that is required to form bonds.
   - The atom that will form the most bonds will be the central atom.

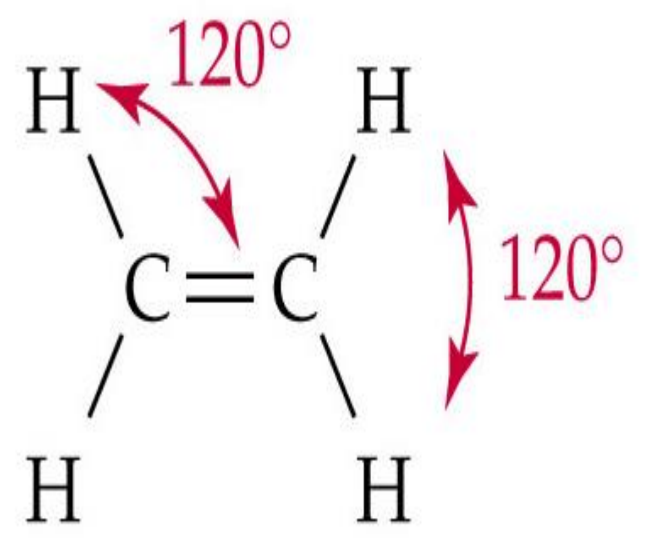
4. Arrange the outer atoms around the central atom so that their single dots are nearest to the central atom.
   - Pair up single electrons between the central atom and the outer atoms to form covalent bonds.

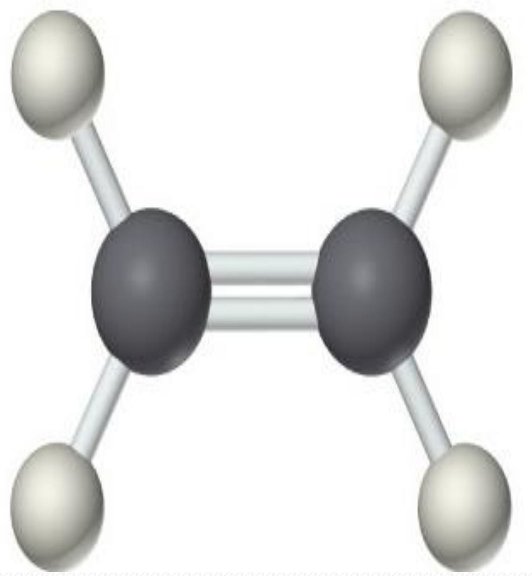
5. Each atom in the structure should have 8 bonding and non-bonding electrons around it (2 for hydrogen).

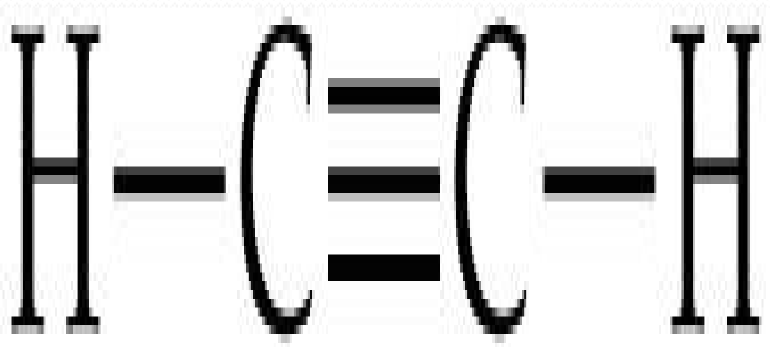


Multiple Bonds

- Sometimes atoms share more than one pair of electrons
- 2 pairs = double bond
- 3 pairs = triple bond

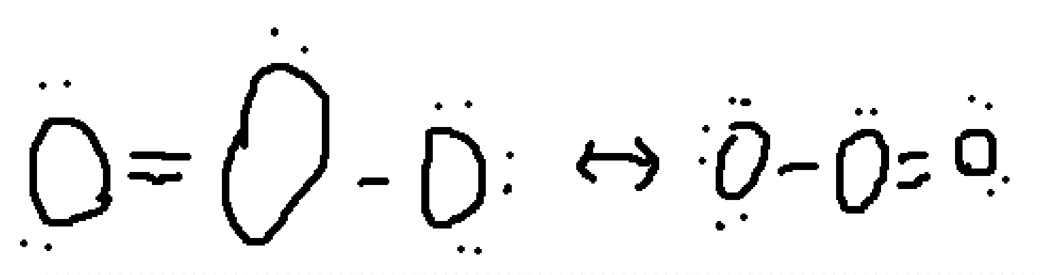


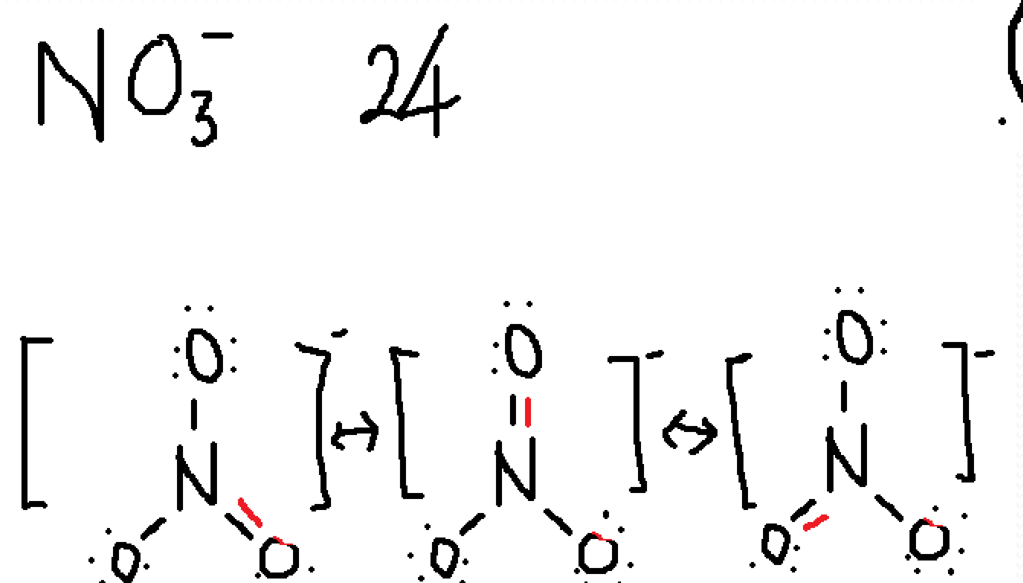




Resonance Structures

- When there are double bonds that can exist in multiple places without changing the structure, all the possibilities must be drawn with double arrows between them.
- None of the structures actually exist, the actual structure is an average of all the resonance structures.
- O_3
- NO_3^-





Try it!

- 1.

(a) H_2S	(b) CCl_4	(c) NCl_3
(d) SO_4^{2-}	(e) HSO_4^-	(f) NO_2^-
(g) CO	(h) PO_4^{3-}	

- 2.

(a) ClO_4^-	(c) CN^-	(e) HCO_3^-	(g) NO^+
(b) CO_3^{2-}	(d) H_3O^+	(f) HNO_3	