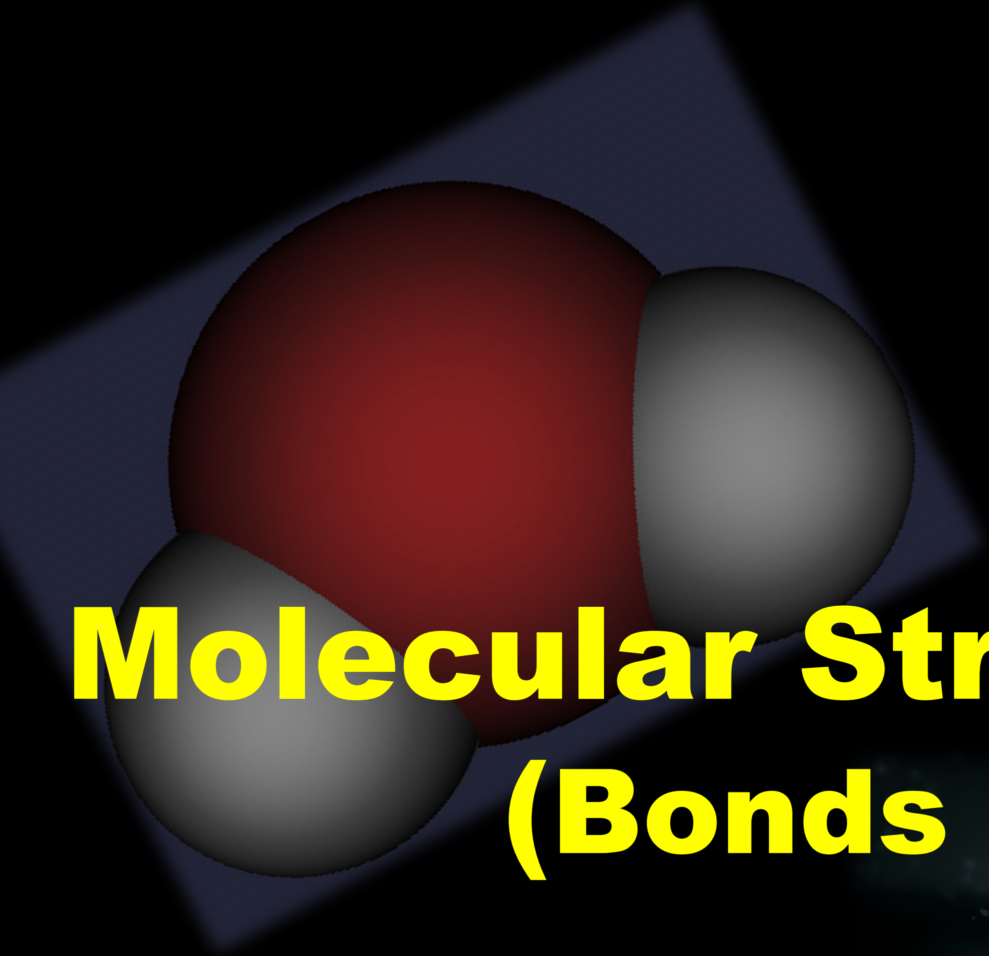
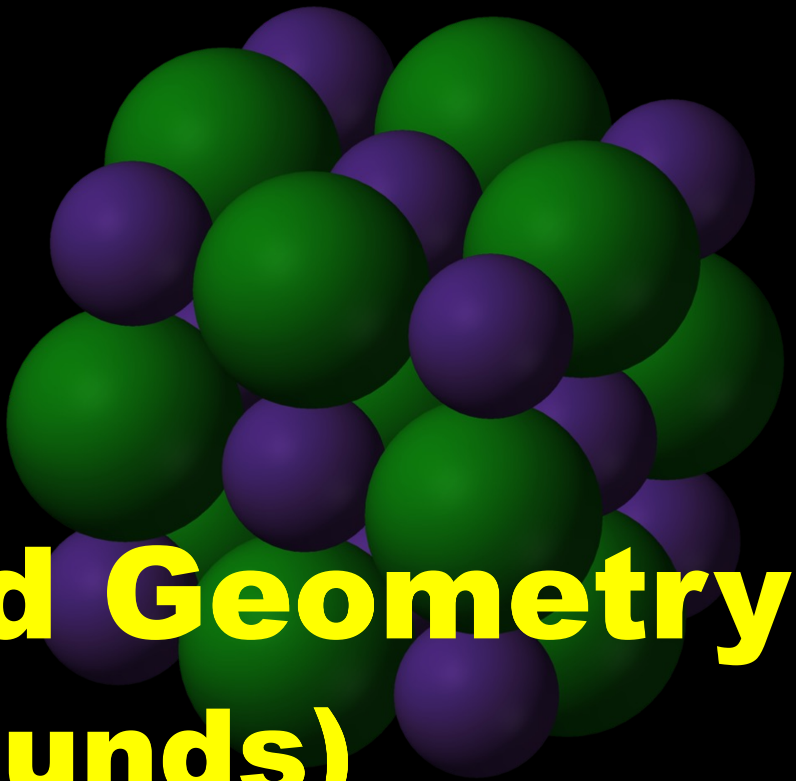



Molecular Structure and Geometry
(Bonds and Compounds)

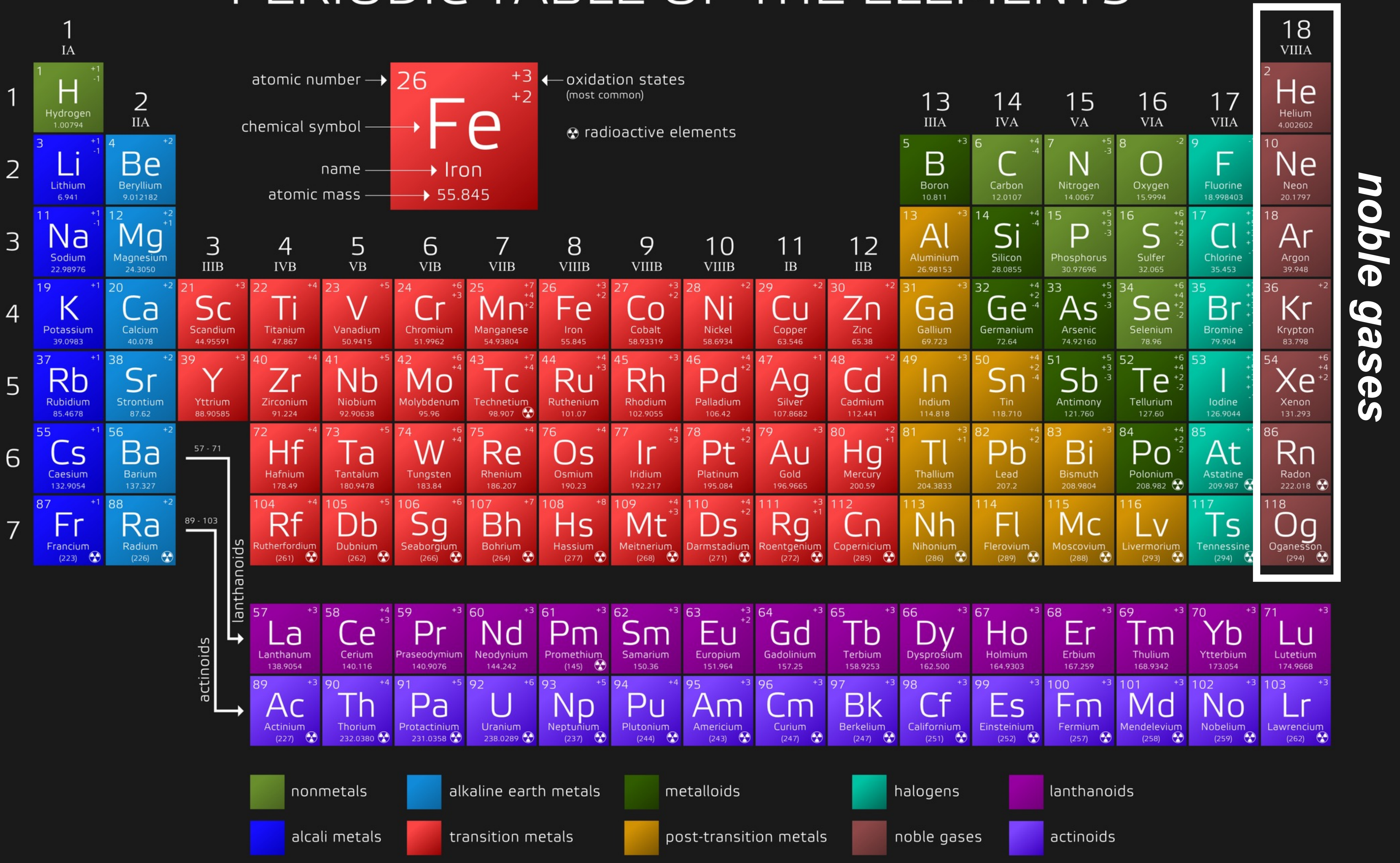
The Octet Rule

Every element wants to "feel" like a ***noble gas***. The noble gases are the elements in Column (or "group") VIII of the periodic table. Group VIII is also called Group 18, if you're including the ***d-block*** in your count.

The ***noble gases*** all have eight ***valence electrons***. Thus, all elements want to "feel" like a ***noble gas*** by either giving away, stealing, or sharing electrons so that they have –or "feel" like they have– eight ***valence electrons***. This is called the ***octet rule***.

Thus, the way elements behave is often driven by the ***octet rule***: that is, their drive to attain eight ***valence electrons***.

PERIODIC TABLE OF THE ELEMENTS



Ions and Shifting Columns

When atoms gain an electron, they become negatively-charged. When atoms lose an electron, they become positively-charged. Atoms that have charges on them are called ***ions***, with positively-charged atoms being called ***cations*** and negatively-charged atoms being called ***anions***.

Each time an element gains one electron, it moves one column to the right on the periodic table, in terms of how it "feels."

Each time an element loses one electron, it moves one column to the left on the periodic table, in terms of how it "feels".

Elements gain, lose, or share electrons –in order to obtain a full ***octet***– by bonding with each other.

Back to Compounds

In an earlier video, we also learned that ***compounds*** are molecules comprised of two or more *different* elements bonded together. Examples include H_2O and CO_2. I'll now teach you about two different classes of compounds:

Compounds

Ionic compounds: metal + nonmetal

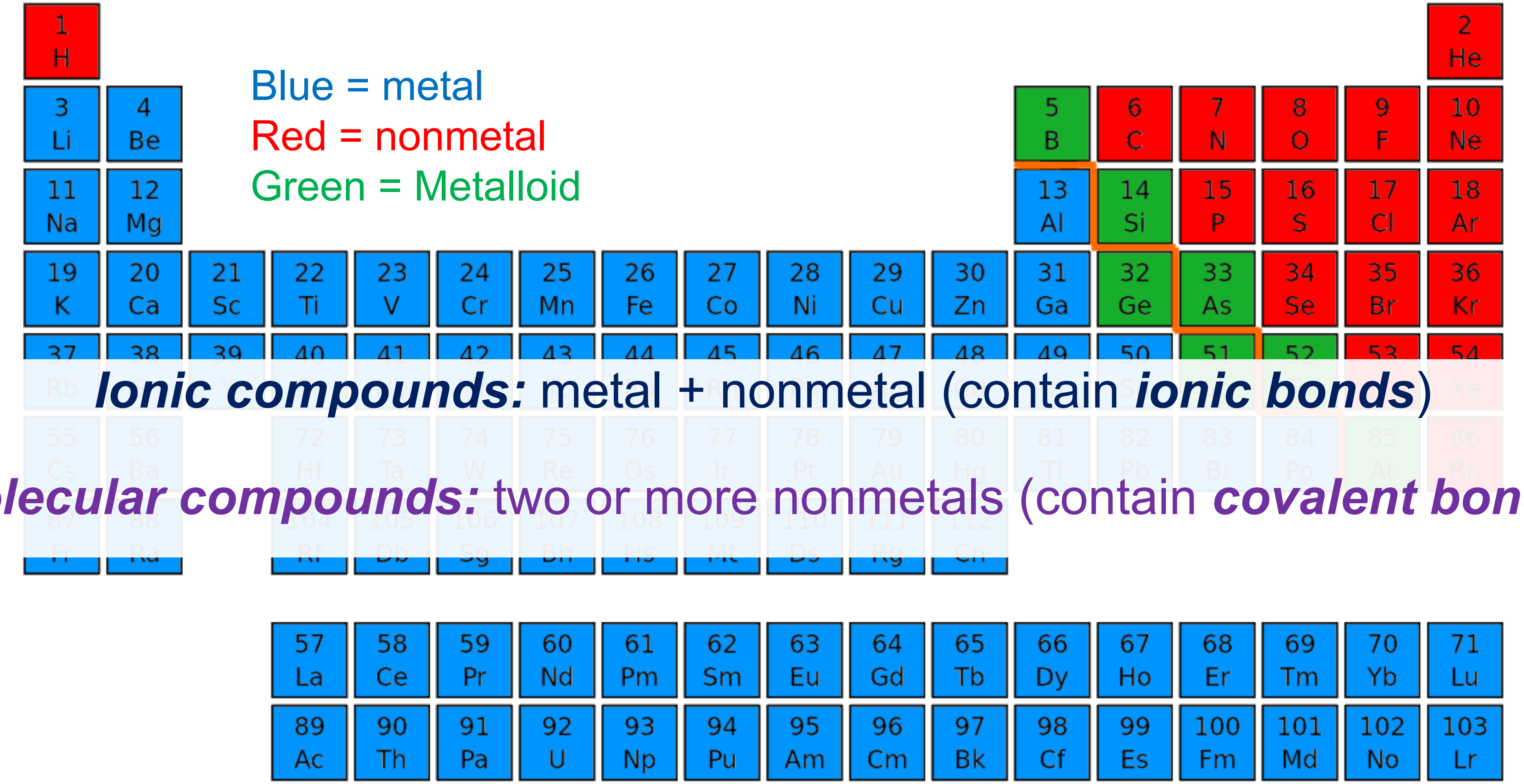
Examples: $NaCl$, $CuCl_2$

Molecular compounds: two or more nonmetals

Examples: H_2O, CO_2

Metals and Nonmetals

And what do I mean when I say ***metals*** and ***nonmetals?***



Ionic compounds: metal + nonmetal (contain ***ionic bonds***)

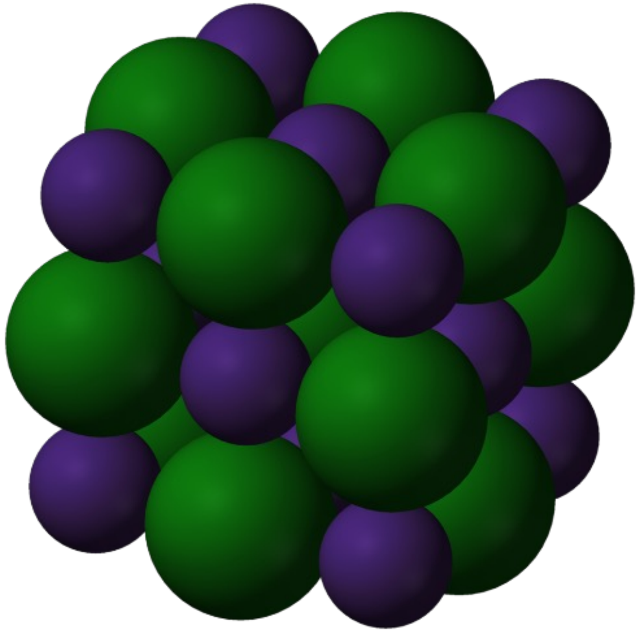
Molecular compounds: two or more nonmetals (contain ***covalent bonds***)

Ionic Compounds

As I just stated, ***ionic compounds*** are formed when a metal and a nonmetal bond with each other. When ***ionic compounds*** form, the ***metal*** pretty much "gives away" its electron to the ***nonmetal***. Thus, the metal becomes a ***cation*** (positively-charged), while the nonmetal becomes an ***anion*** (negatively-charged):

$$Na \; + \; Cl_2 \; \rightarrow \; NaCl \; = \; Na^+ \; Cl^-$$

Once the ***cation*** and ***anion*** form, however, they don't just go their separate ways. They instead cluster together to make an ionic solid.



NaCl

Ionic Exceptions

Once again, ***ionic compounds*** are usually formed when a metal and a nonmetal bond with each other. There is, however, one class of ionic compounds that are made of *all nonmetals*.

When ***polyatomic ions*** bond together, they also make ***ionic compounds***, even though they do NOT contain any metals. For example, ammonium chloride, NH_4Cl, is ionic, even though it doesn't contain any metals.

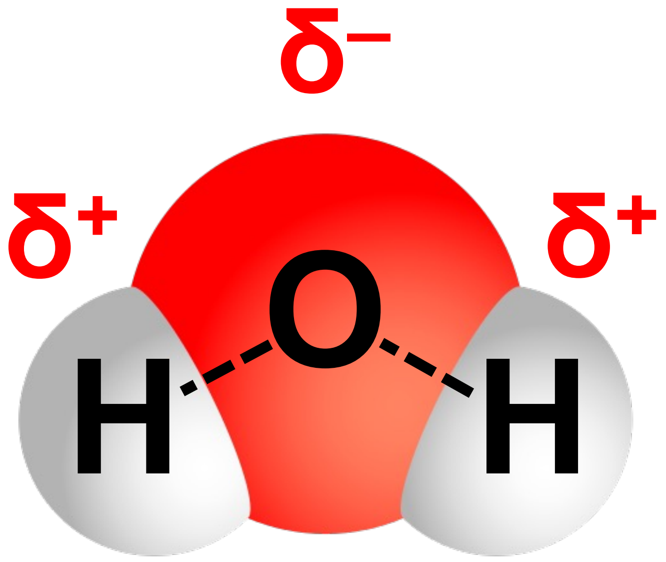
This is because NH_4^+ (ammonium) is a ***cation***, and Cl^- (chloride) is an ***anion***. Again, when ***polyatomic ions*** bond together, they also make ionic compounds, even though they do NOT contain any metals.

Polyatomic ions you should <u>memorize</u> for the EXAM.

+1 CHARGE		-1 CHARGE		-2 CHARGE		-3 CHARGE	
ion	**name**	**ion**	**name**	**ion**	**name**	**ion**	**name**
NH_4^+	ammonium	$H_2PO_3^-$	dihydrogen phosphite	HPO_3^{2-}	hydrogen phosphite	PO_4^{3-}	phosphate
H_3O^+	hydronium	$H_2PO_4^-$	dihydrogen phosphate	HPO_4^{2-}	hydrogen phosphate	AsO_3^{3-}	arsenite
Hg_2^{2+}	mercury(I)	HCO_3^-	hydrogen carbonate	CO_3^{2-}	carbonate	AsO_4^{3-}	arsenate
		HSO_3^-	hydrogen sulfite	SO_3^{2-}	sulfite	N^{3-}	nitride
		HSO_4^-	hydrogen sulfate	SO_4^{2-}	sulfate		
		NO_2^-	nitrite	$S_2O_3^{2-}$	thiosulfate		
		NO_3^-	nitrate	$C_2O_4^{2-}$	oxalate		
		OH^-	hydroxide	CrO_4^{2-}	chromate		
		CH_3COO^-	acetate	$Cr_2O_7^{2-}$	dichromate		
		CrO_2^-	chromite	O_2^{2-}	peroxide		
		CN^-	cyanide	S_2^{2-}	disulfide		
		CNO^-	cyanate	O^{2-}	Oxide		
		CNS^-	thiocyanate	S^{2-}	Sulfide		
		MnO_4^-	permanganate				
		ClO^-	hypochlorite				
		ClO_2^-	chlorite				
		ClO_3^-	chlorate				
		ClO_4^-	perchlorate				
		BrO^-	hypobromite				
		BrO_2^-	bromite				
		BrO_3^-	bromate				
		BrO_4^-	perbromate				
		IO^-	hypoiodite				
		IO_2^-	iodite				
		IO_3^-	iodate				
		IO_4^-	periodate				
		N_3^-	azide				

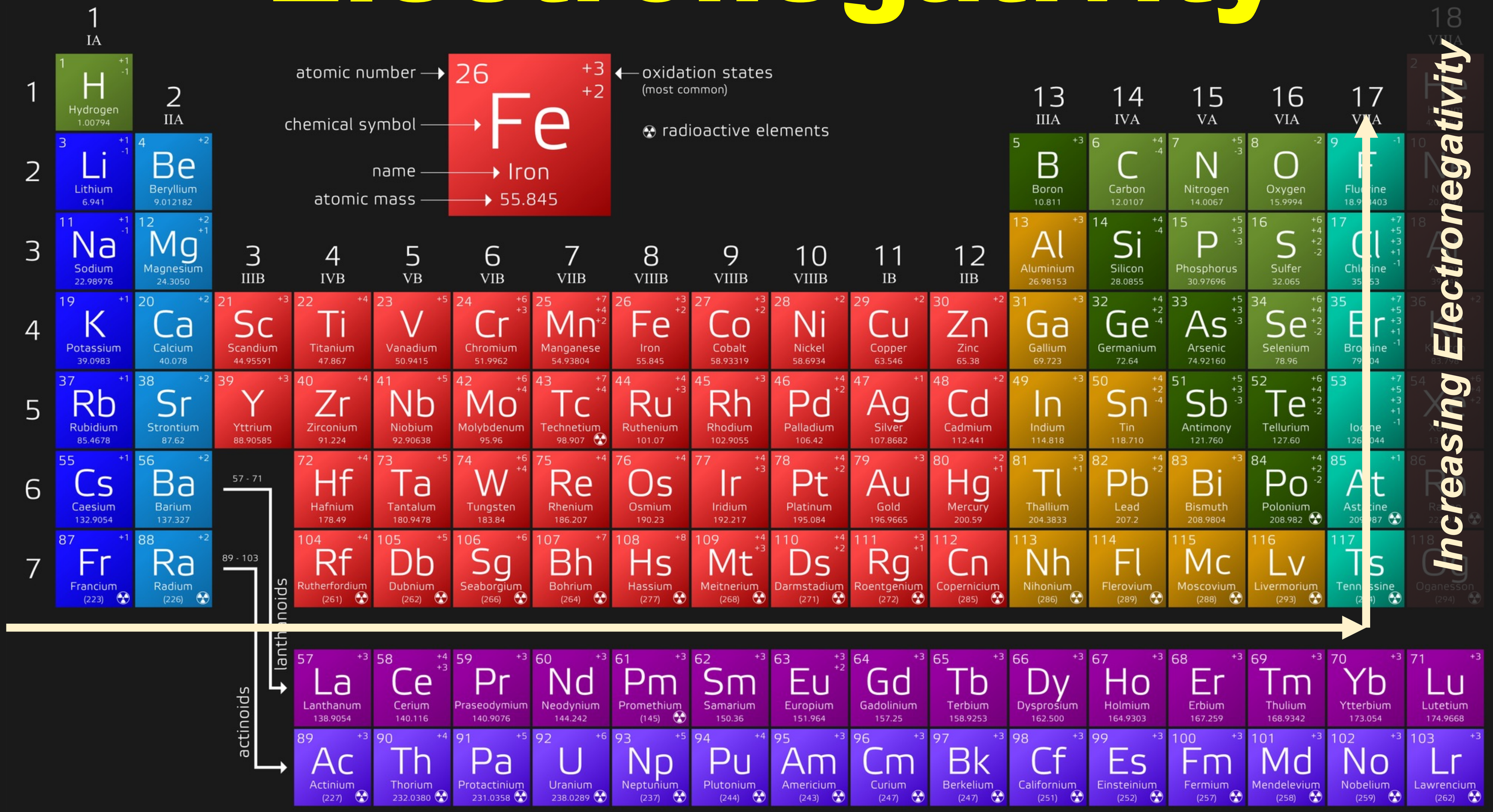
Molecular Compounds

Molecular compounds are formed when two or more nonmetals bond with each other. Bonds between nonmetals are called ***covalent bonds***. In ***covalent bonds***, the atoms *share* their electrons, so they do NOT have *true* positive or negative charges (***cations*** or ***anions***). Instead, they can have *partial* positive and negative charges, if they share their electrons unevenly.



If they share their electrons evenly, then they do NOT have partial charges.

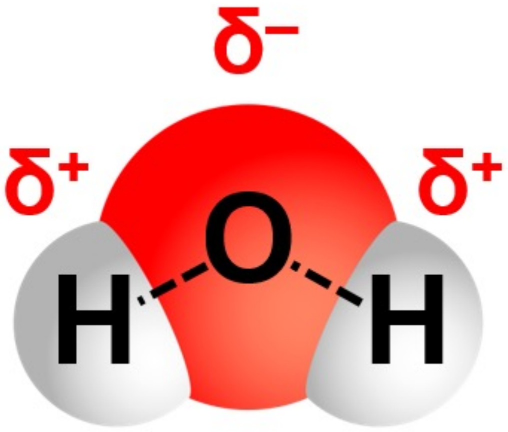
Electronegativity



An atom's ***electronegativity*** is its "thirst" for electrons. ***Electronegativity*** increases as you go up and to the right on the periodic table.

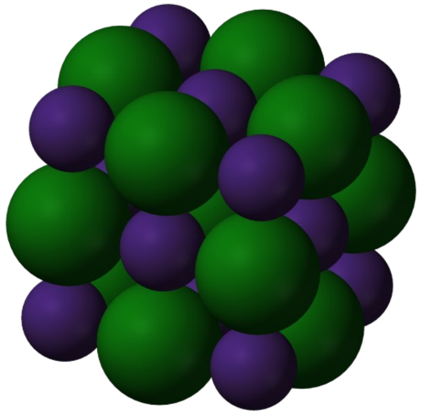
Polar Covalents

Bonds in ***molecular compounds*** that have an uneven sharing of electrons (which is caused by a significant ***electronegativity*** difference between the bonded atoms) are called ***polar covalent bonds***. They're kind of like ***ionic bonds***, except they don't have <u>*true*</u> cations or anions. They just have <u>partial</u> positive or negative charges.



H_2O

versus

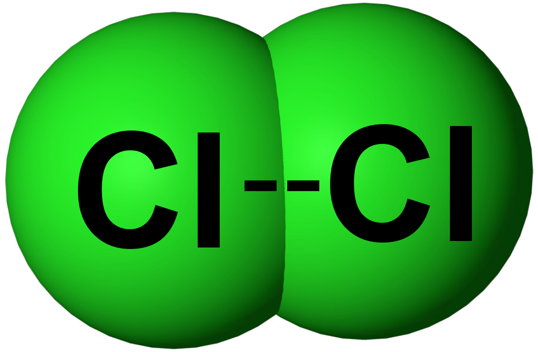


NaCl

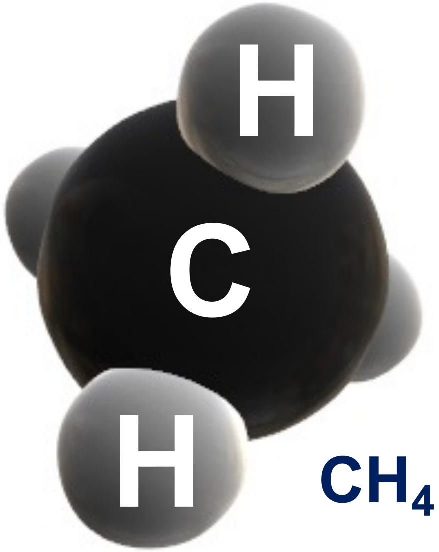
So you can think of a ***polar covalent bond*** as kind of having "partial" ionic character.

Nonpolar Covalents

This contrasts with ***nonpolar covalent bonds***, which occur when ***nonmetal*** bonded atoms do NOT have a significant ***electronegativity*** difference between them. ***Nonpolar covalent molecules*** do <u>NOT</u> have partial positive or negative charges.

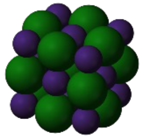
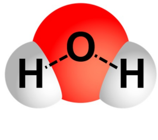
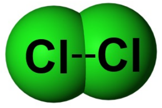
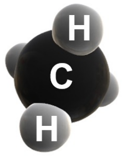
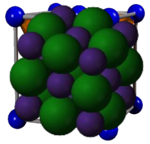
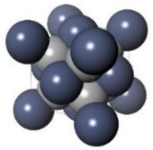


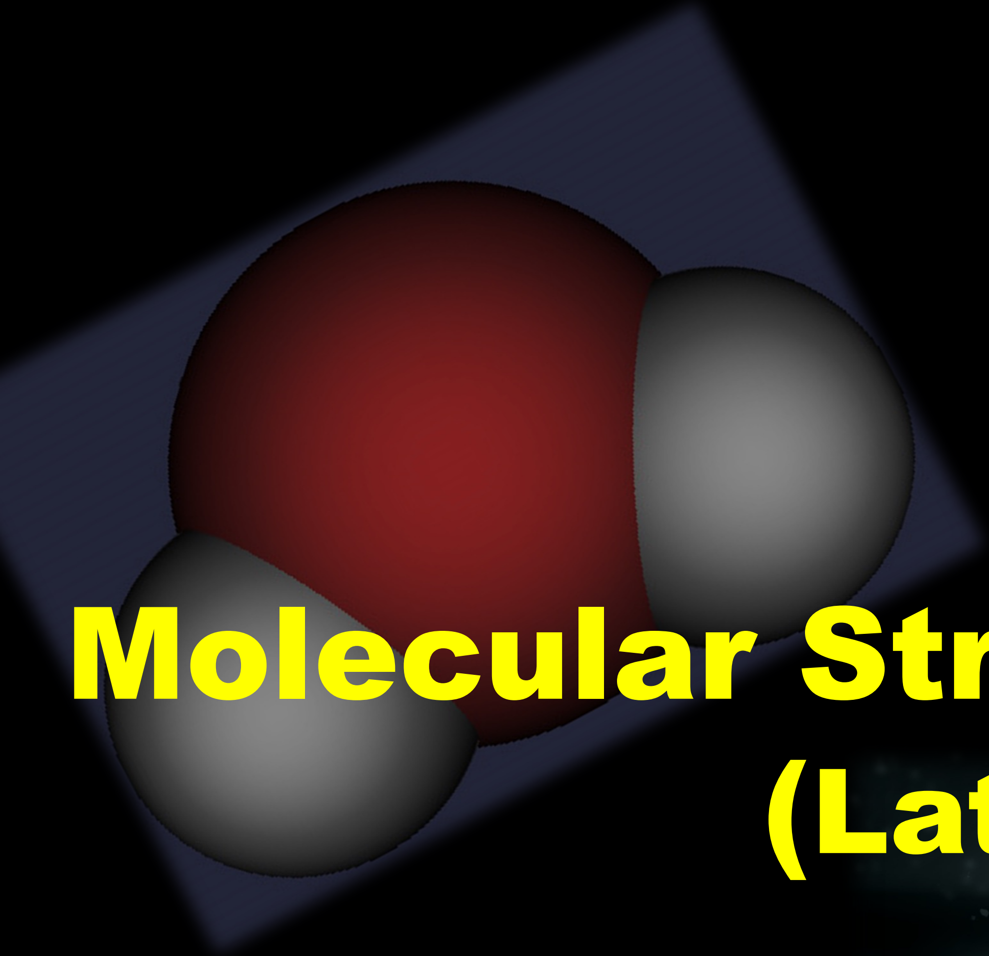
Cl_2

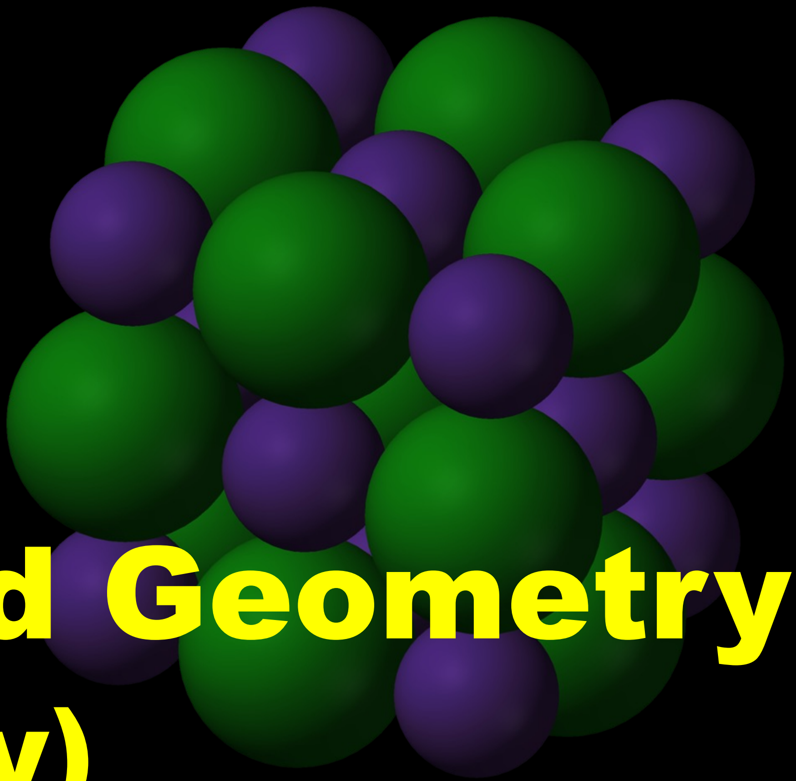


CH_4

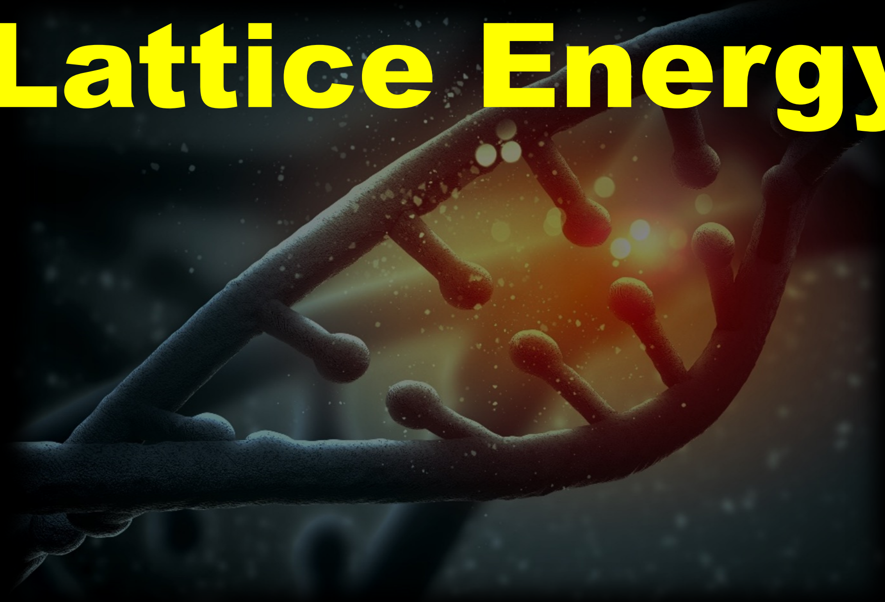
Different Compound Types

Type	Properties	What keeps them together?	Examples
 Ionic	high melting points, high boiling points, brittle, hard	Ionic interactions (lattice energy)	NaCl, MgO
Molecular	low melting points, do not conduct electricity	Intermolecular Forces	  
 Network Covalent	high melting points, high boiling points, hard, do not conduct electricity	Network of covalent bonds	C (diamond, graphite) SiO_2 (quartz)
 Metallic	variable hardness and melting points, conduct electricity, conduct heat, lustrous (shiny), malleable, ductile	Metallic Bonding	Fe, Mg

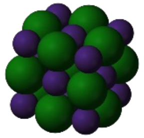
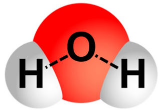
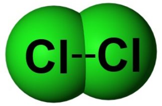
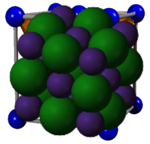
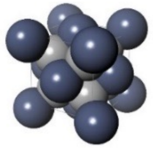




Molecular Structure and Geometry
(Lattice Energy)



Different Compound Types

Type	Properties	What keeps them together?	Examples
 Ionic	high melting points, high boiling points, brittle, hard	Ionic interactions (lattice energy)	NaCl, MgO
Molecular	low melting points, do not conduct electricity	Intermolecular Forces	 
 Network Covalent	high melting points, high boiling points, hard, do not conduct electricity	Network of covalent bonds	C (diamond, graphite) SiO_2 (quartz)
 Metallic	variable hardness and melting points, conduct electricity, conduct heat, lustrous (shiny), malleable, ductile	Metallic Bonding	Fe, Mg

What is Lattice Energy?

Lattice energy is the energy required to completely separate an ionic compound's ***cations*** from its ***anions***.

$$NaCl_{(s)} \rightarrow Na^{+}_{(g)} + Cl^{-}_{(g)}$$

What is Lattice Energy?

Lattice energy is the energy required to completely separate an ionic compound's ***cations*** from its ***anions***.

$$\textbf{\textit{Lattice energy}} \approx \frac{\textbf{(cation charge)} \times \textbf{(anion charge)}}{\textbf{Bond distance}}$$

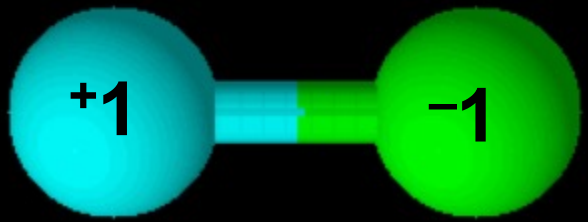
Bigger charges = larger lattice energy

Shorter bond distance = larger lattice energy

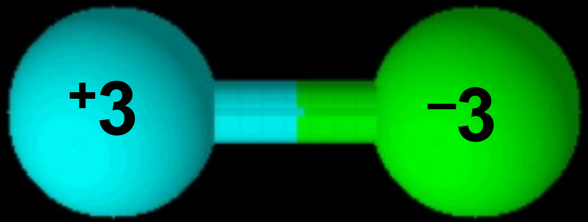
Bigger Charges = Larger Lattice Energy

$$\textit{Lattice energy} \approx \frac{\text{(cation charge)} \times \text{(anion charge)}}{\text{Bond distance}}$$

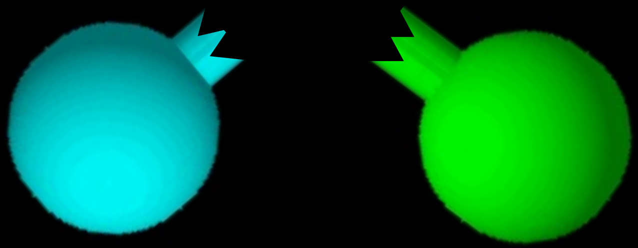
Scenario A:



Scenario B:



In which scenario is it harder to break apart the atoms?

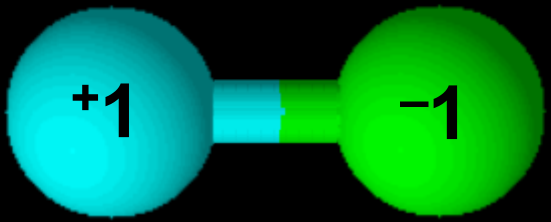


bigger charges = larger lattice energy

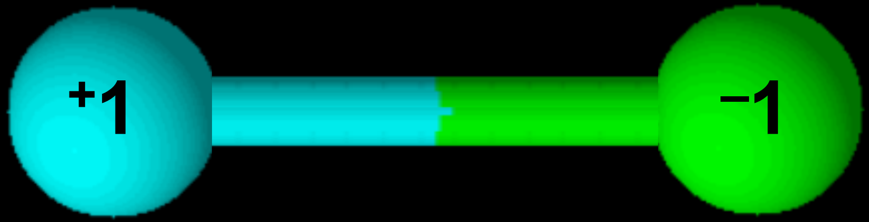
Shorter Bond = Larger Lattice Energy

$$\textit{Lattice energy} \approx \frac{\text{(cation charge)} \times \text{(anion charge)}}{\text{Bond distance}}$$

Scenario A:



Scenario B:



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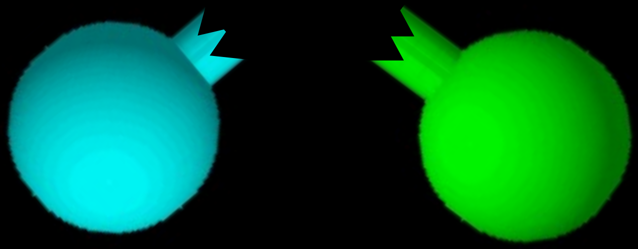
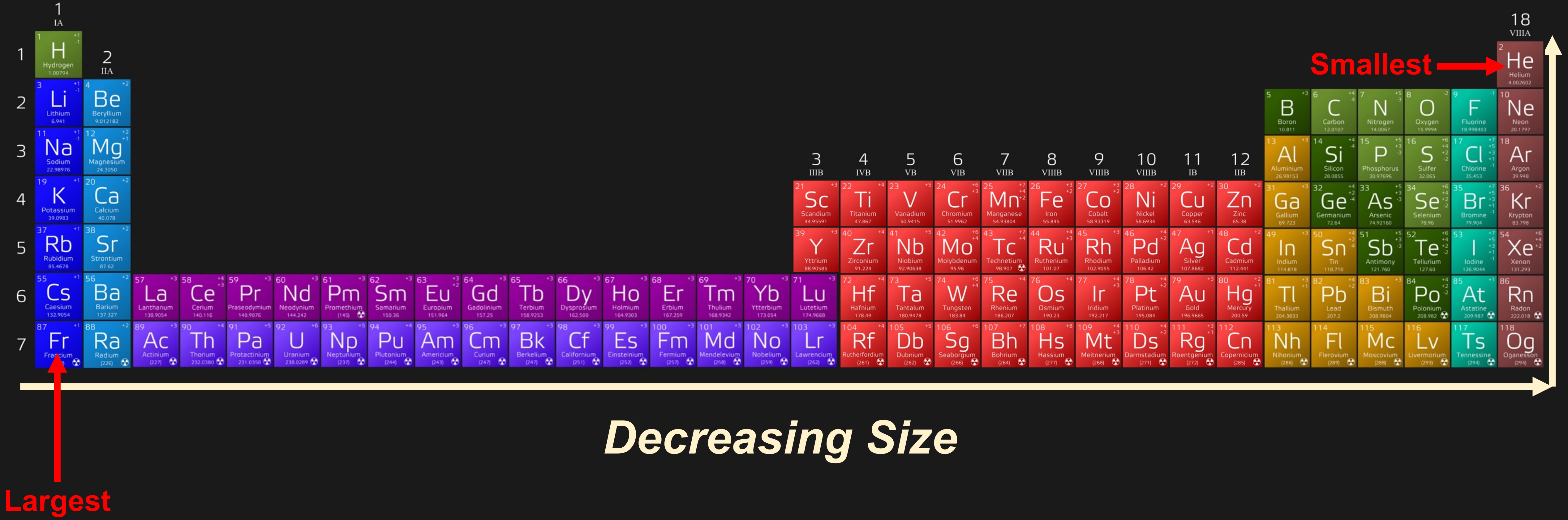


shorter bond distance
= larger lattice energy

Elements' Sizes

In general, atoms' sizes get ***smaller*** as you go to the right across a row or up a column on the periodic table. You'll notice that this follows the same trend as ***electronegativity***, except noble gases are included.



Example Questions

Which of the following would have the largest ***lattice energy?*** Why?

A. Li-F

B. Li-Cl

C. Li-Br

D. Li-I

Example Questions

Which of the following would have the largest ***lattice energy?*** Why?

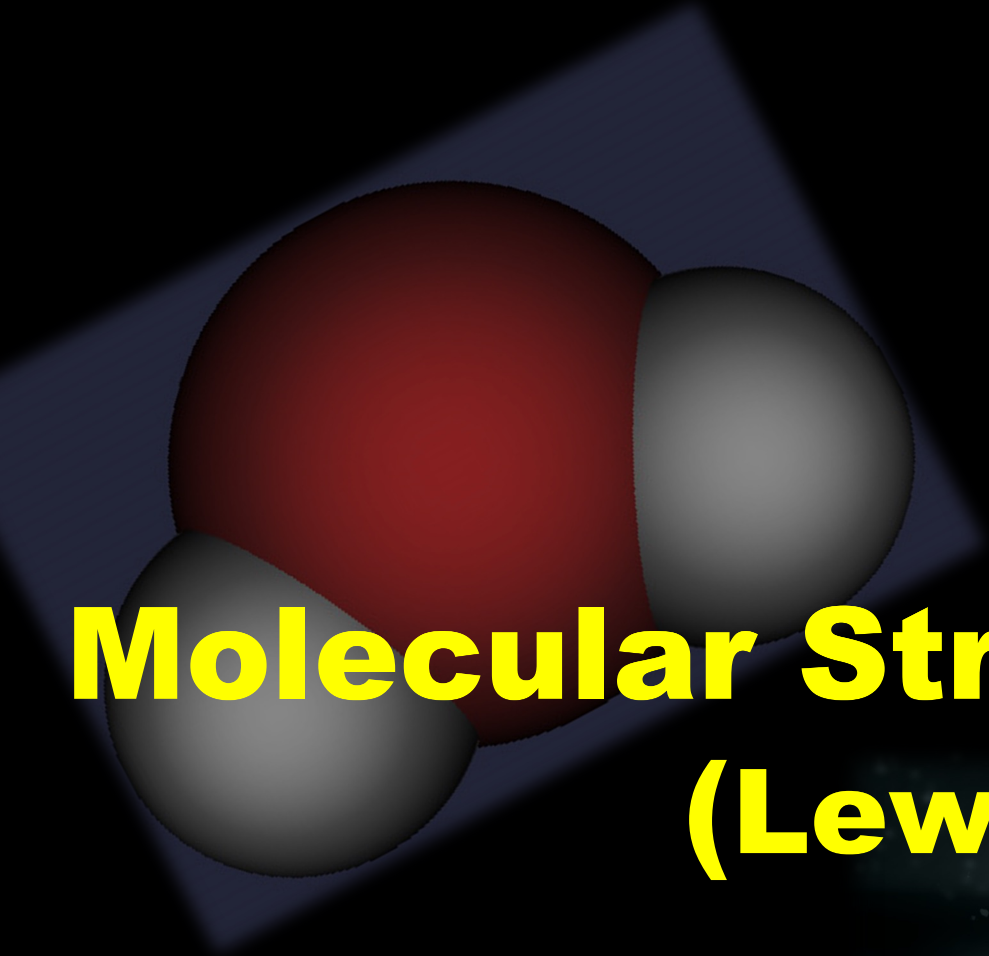
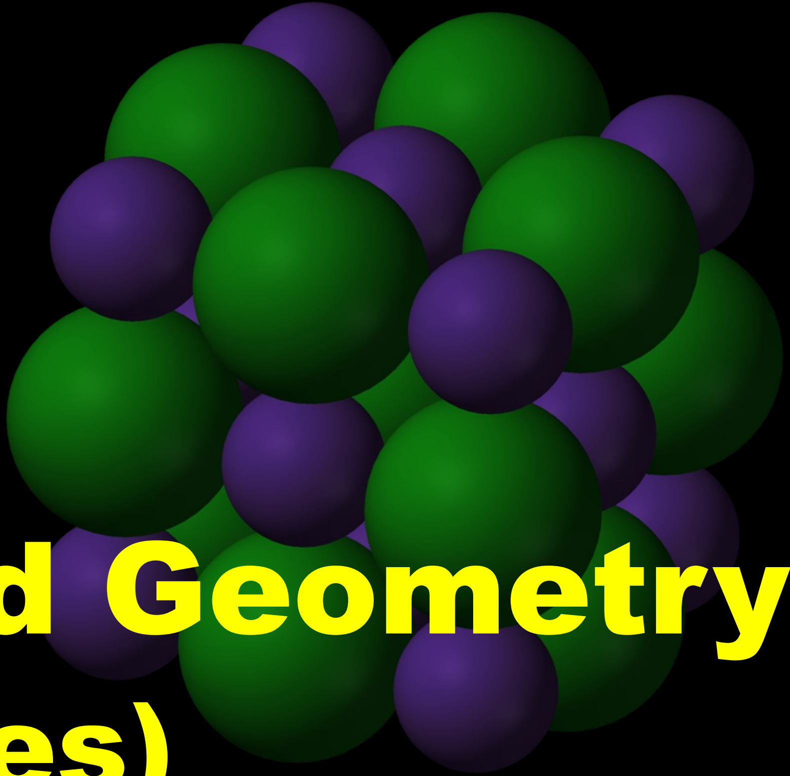
A. Li-F

B. Na-F

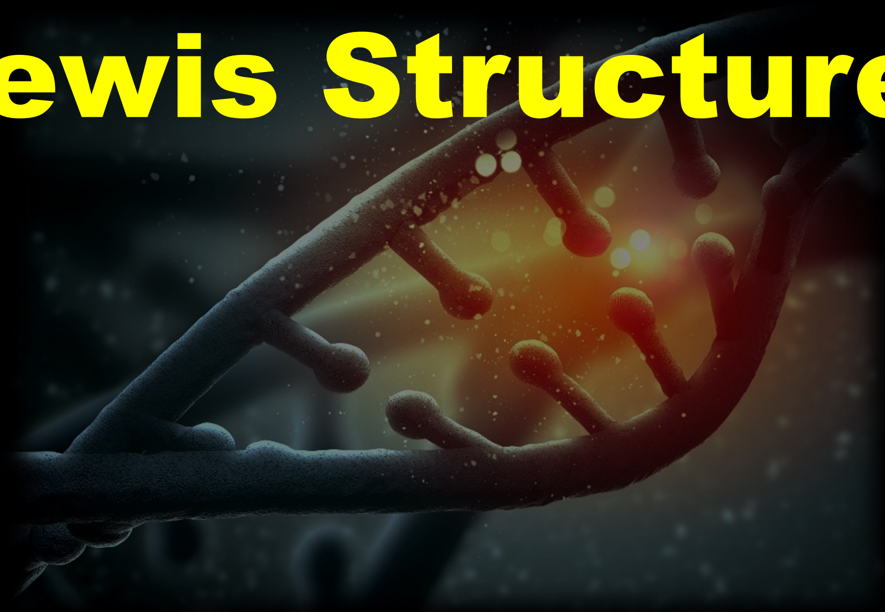
C. $MgCl_2$

D. ScN

E. CaO

Molecular Structure and Geometry (Lewis Structures)



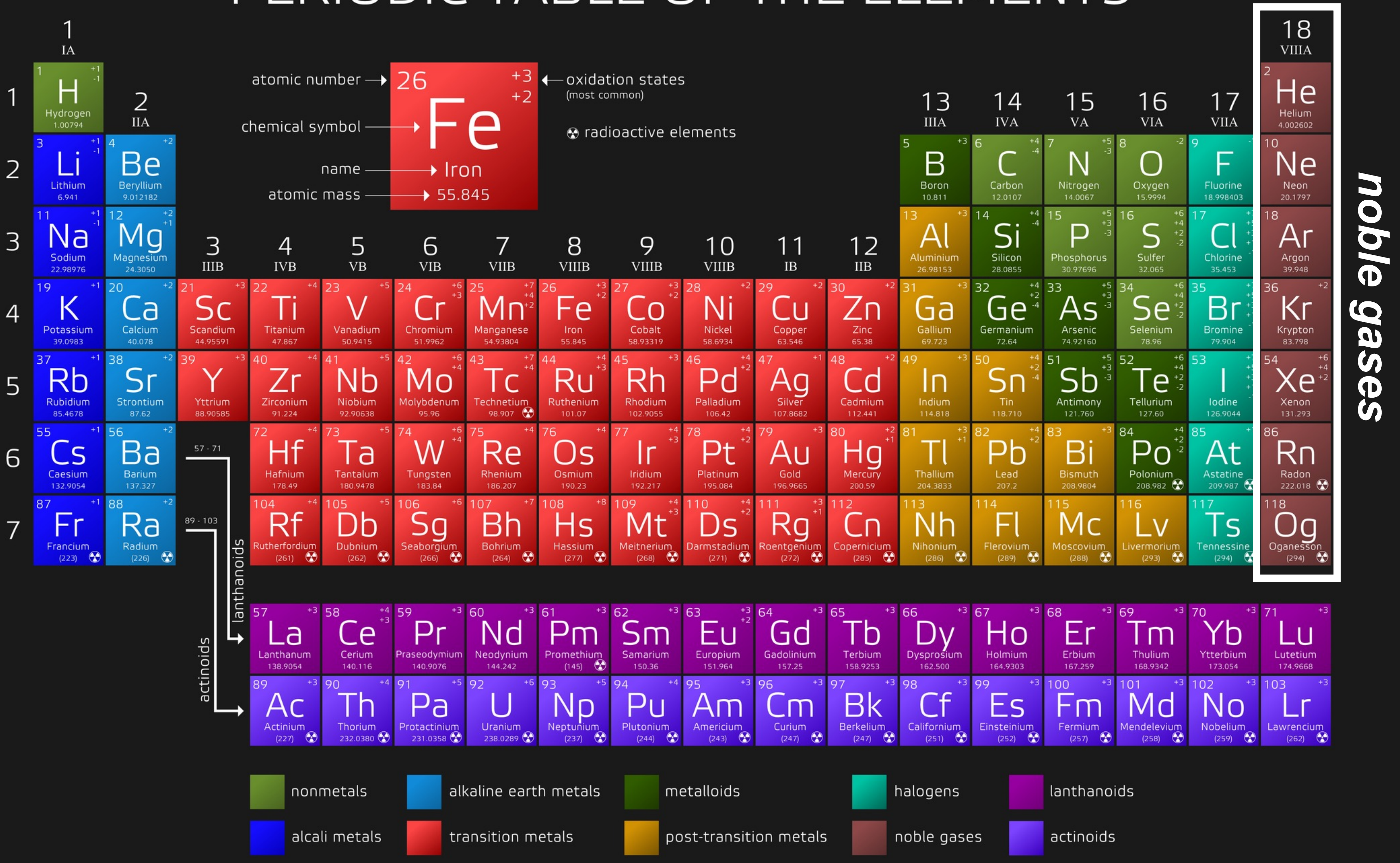
The Octet Rule

Every element wants to "feel" like a ***noble gas***, which are the elements in Group VIII (or Group 18, if you count the ***d-block***) of the periodic table.

The ***noble gases*** all have eight ***valence electrons***. Thus, all elements want to "feel" like a ***noble gas*** by either giving, stealing, or sharing electrons so that they have –or "feel" like they have– eight ***valence electrons***. This is called the ***octet rule***.

Thus, elements' chemical properties are significantly driven by the ***octet rule***: their determination to attain eight ***valence electrons***.

PERIODIC TABLE OF THE ELEMENTS



Ions and Shifting Columns

Each time an element gains one electron through stealing or sharing, it moves one column to the right on the periodic table, in terms of how it "feels."

Each time an element loses one electron through stealing or sharing, it moves one column to the left on the periodic table, in terms of how it "feels".

Elements gain, lose, or share electrons –in order to obtain a full ***octet***– by bonding with each other.

How to Draw Lewis Structures

1. Count up the valence electrons for every atom in your formula. Add one electron to your count for each negative charge. Subtract one electron from your count for each positive charge. (Note: an atom's number valence electrons is usually the same as its column number on the periodic table, except that we *exclude* the columns in the ***d-*** and ***f-blocks***. For elements that are *in* the ***d-*** or ***f-blocks***, we just count the electrons in their outermost ***s-*** or ***p-***columns as being valence electrons.)

2. Write down the symbols for all the atoms in your molecule. Usually, the least electronegative atom goes in the center. This is typically the first atom in your chemical formula, except for H. (H is NEVER in the middle.) Connect each atom a single bond, remembering that each single bond represents two electrons.

3. Use your remaining electrons to complete everyone's octet by adding ***lone-pair electrons*** (also called ***nonbonding pairs***) around each atom, leaving the central atom for last. (Note: hydrogen only needs a two-tet, not an octet.)

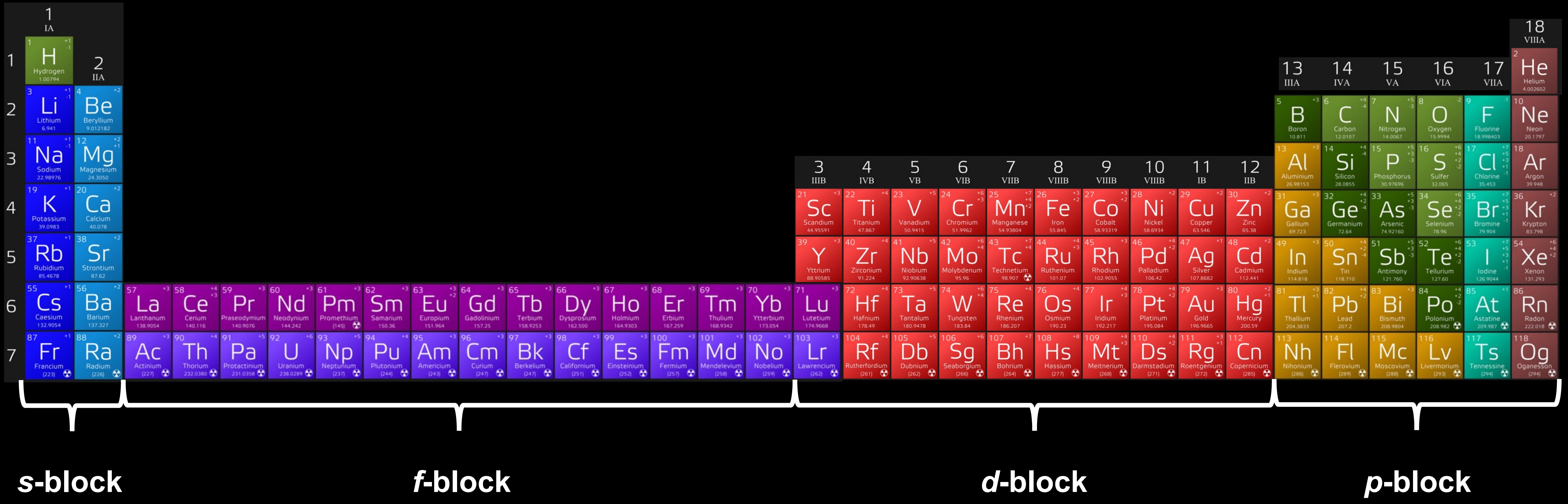
How to Draw Lewis Structures

4. Place any leftover electrons on the central atom, even if doing so violates the octet rule. (<u>Note</u>: elements on the third row and below on the periodic table CAN have more than eight electrons.)

5. If there are not enough electrons to give the central atom an octet, form multiple bonds (like double or triple bonds) to complete the central atom's octet.

How Many Valence Electrons?

- Usually the same as its column number on the periodic table



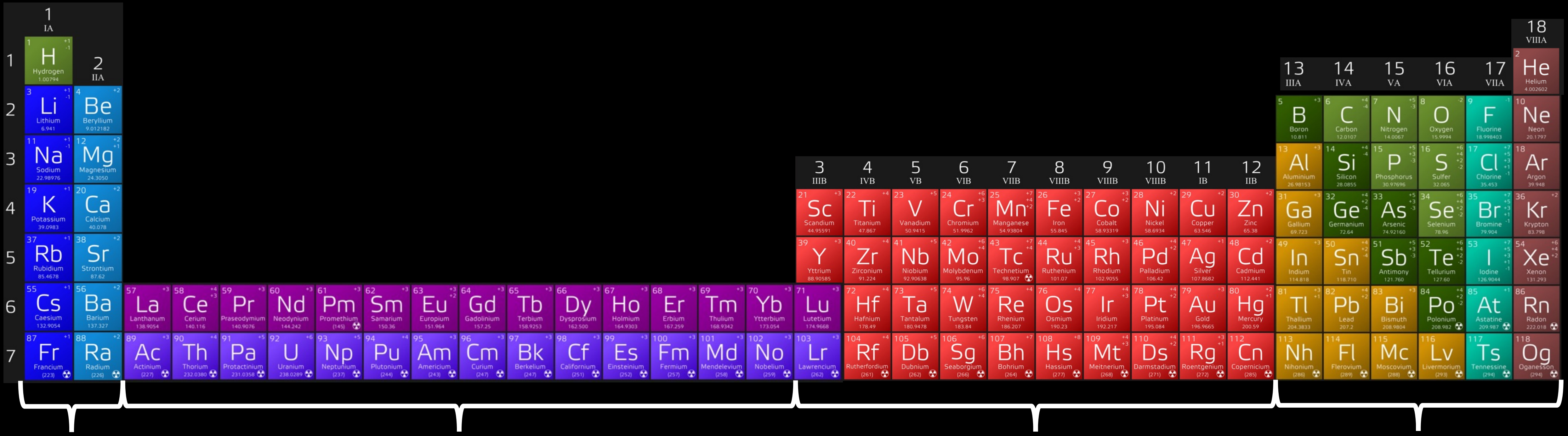
How Many Valence Electrons?

"We do not consider the electrons in COMPLETELY FILLED **d** *or* **f** *subshells to be valence electrons,* and *for transition elements, we do not consider the electrons in a completely filled* **f** *subshell to be valence electrons."*

– Brown, T. L.; LeMay, Jr.; H. E.; Bursten, B. E.; Murphy, C. J.; Woodward, P. M.; Stoltzfus, M. W. Electronic Structure of Atoms. In *Chemistry: the Central Science*, 13th edition; Pearson: River, NJ; 2015; p 244.

How Many Valence Electrons?

- Usually the same as its column number on the periodic table
- If ***d-block*** is completely filled, ***d-block*** electrons do NOT count.



***s*-block** ***f*-block** ***d*-block** ***p*-block**

How Many Valence Electrons?

Best Answer: I have no idea what these guys are smoking or if they forgot what the term VALENCE ELECTRON means, but a valence electron is the outermost electron. Yes, the electron configuration of cobalt is what you wrote, BUT the OUTERMOST shell is the 4s shell. Atoms fill electrons in increasing energy, thus it requires less energy to fill the 4s orbital than to fill the 3d orbital. There are TWO valence electrons in cobalt. That is why when cobalt forms an ion it forms Co+2 most often.

Cobalt has two valence electrons.

The electron configuration or Co is

[Ar] 4s^2 3d^7

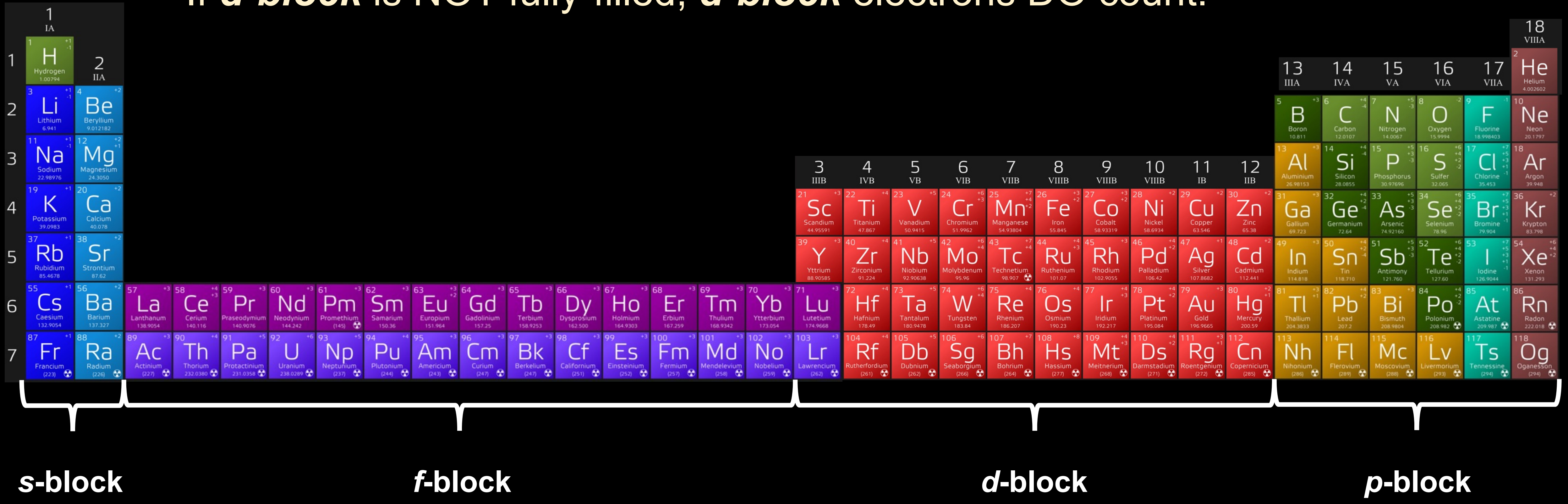
There are 2 in the 4s orbital and 7 in the 3d orbital.

Co has 2 + 7= 9 valence electrons.

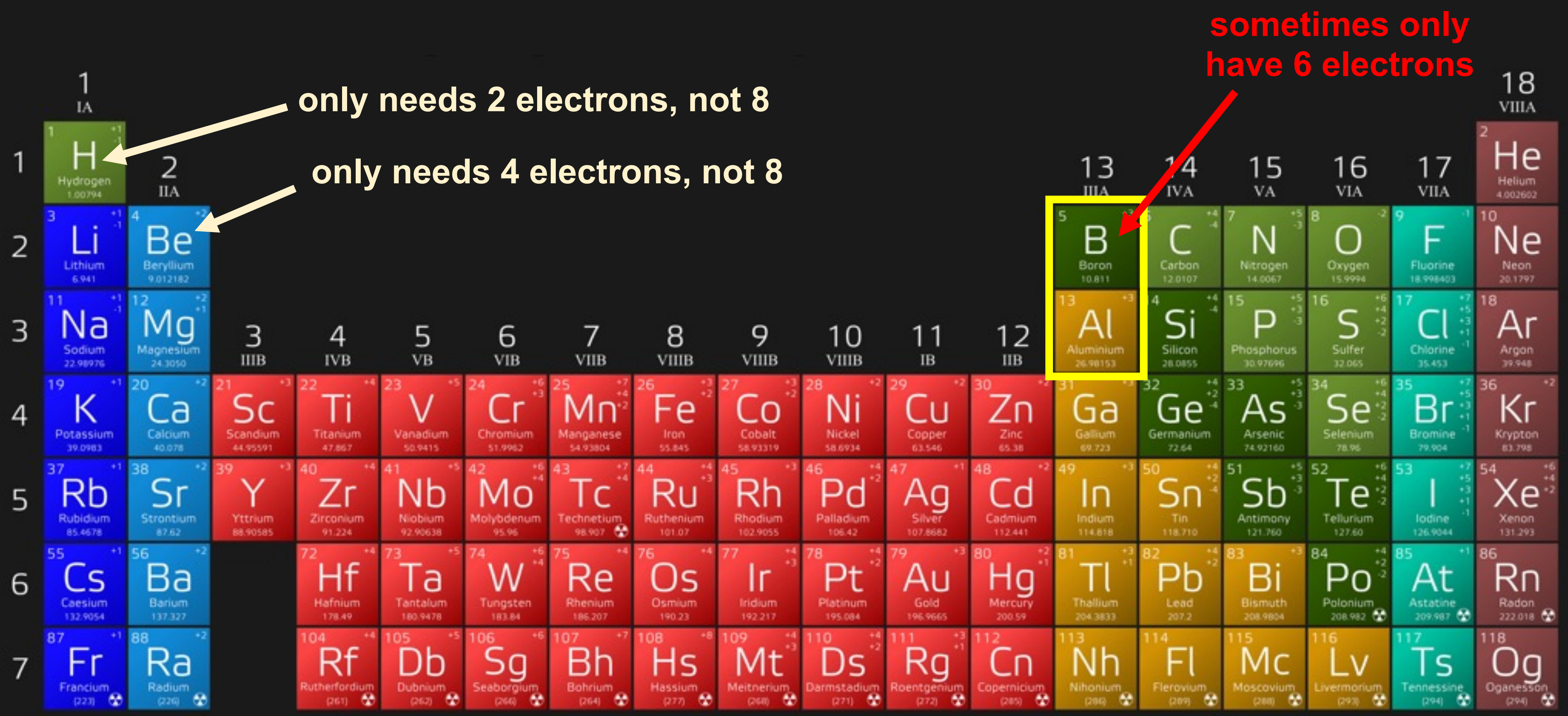
There's 2 or 9 valence electrons, depending on context.

How Many Valence Electrons?

- Usually the same as its column number on the periodic table
- If ***d-block*** is completely filled, ***d-block*** electrons do NOT count.
- If ***d-block*** is NOT fully-filled, ***d-block*** electrons DO count.



Please remember the following exceptions:



Example Questions

Please draw a correct ***Lewis structure*** for each of the following molecules. For examples that have multiple correct ***Lewis structures***, choose the one that represents the most stable ***resonance contributor***.

A. CH_4

B. NH_3

C. N_2O

D. SCl_4

E. NO_3^-

Example Questions

Which of the following molecules contains a triple bond?

I. N_2
II. CH_3CN
III. C_2H_2

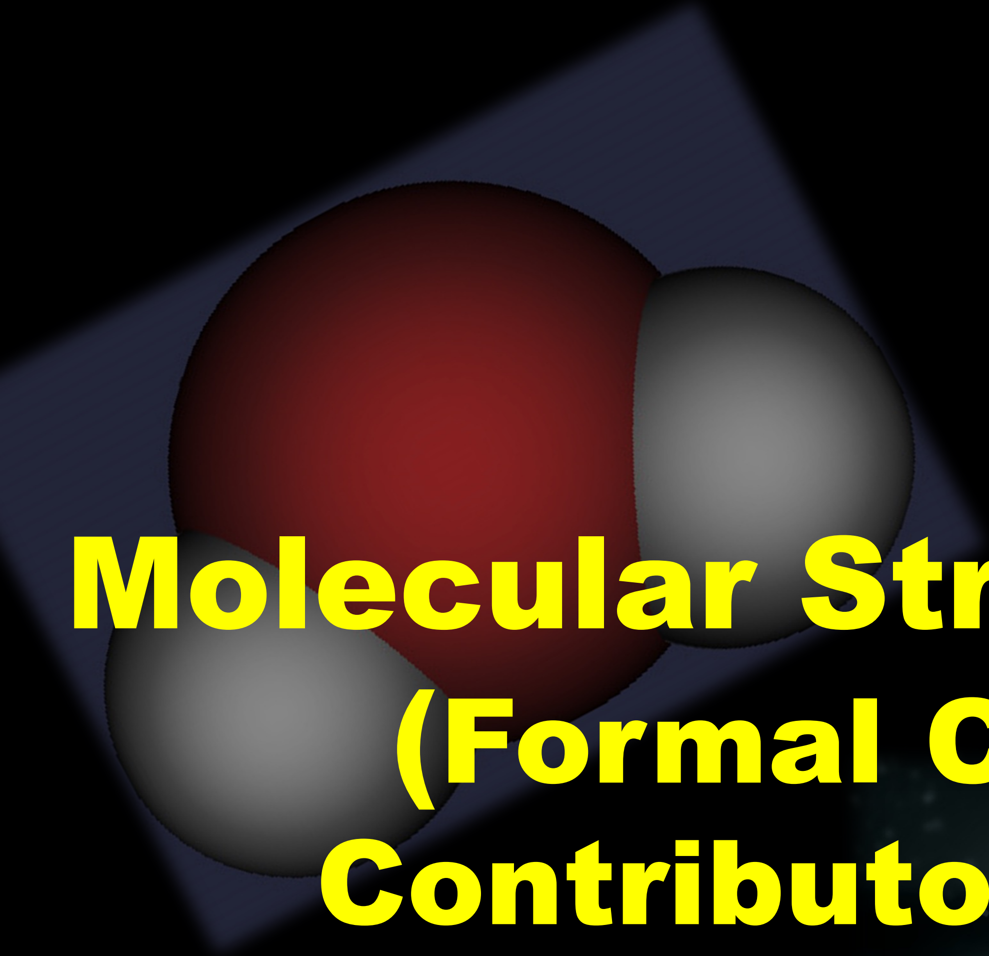
A. I only

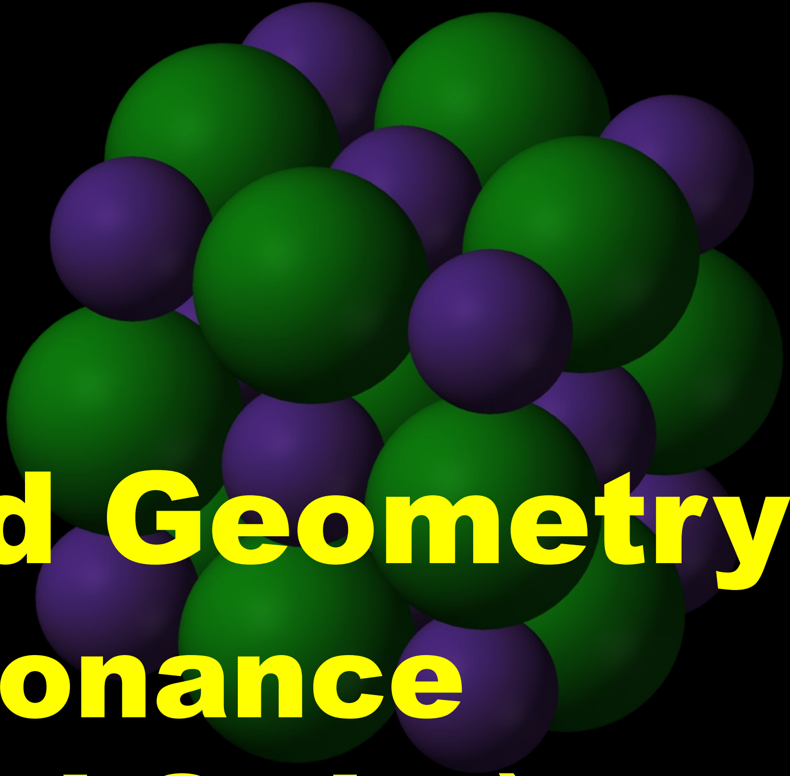
B. I and III only

C. II and III only

D. III only

E. I, II, and III

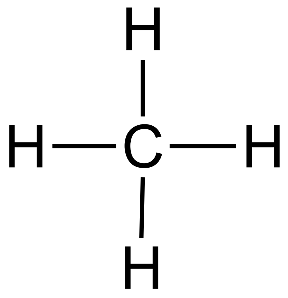
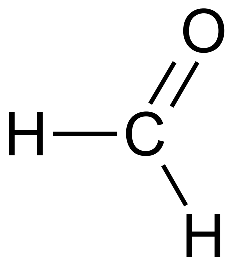




Molecular Structure and Geometry (Formal Charge, Resonance Contributors, and Bond Order)

How to Count σ and π Bonds

Each single bond in a molecule is one ***σ*** (***sigma***) bond. All double bonds contain one ***σ*** and one ***π*** (pi) bond. Triple bonds contain one ***σ*** and two ***π*** bonds:

Molecule	**Bond Count**
	Four ***σ*** bonds
	The C=O bond contains one ***σ*** bond and one ***π*** bond. Each C–H bond is a ***σ*** bond.

How to Count σ and π Bonds

Each single bond in a molecule is one ***σ*** (***sigma***) bond. All double bonds contain one ***σ*** and one ***π*** (pi) bond. Triple bonds contain one ***σ*** and two ***π*** bonds:

Molecule	**Bond Count**
H—C≡C—H	The C≡C bond contains one ***σ*** bond and two ***π*** bonds. Each C–H bond is a ***σ*** bond.

Sigma bonds are stronger than pi bonds. However, because double bonds contain one ***σ*** and one ***π***, double bonds are stronger and shorter than single bonds. By analogy, triple bonds are stronger and shorter than doubles, because they contain one ***σ*** and two ***π*** bonds.

How to Count σ and π Bonds

Each single bond in a molecule is one *σ* (***sigma***) bond. All double bonds contain one *σ* and one *π* (pi) bond. Triple bonds contain one *σ* and two *π* bonds:

Bond gets longer and weaker

Single bond: one ***σ***

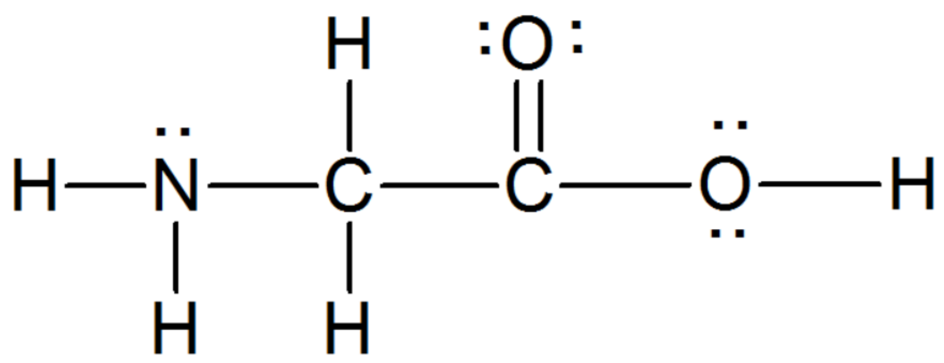
Double bond: one ***σ*** and one ***π***

Triple bond: one ***σ*** and two ***π***

Sigma bonds are stronger than pi bonds. However, because double bonds contain one *σ* and one *π*, double bonds are stronger and shorter than single bonds. By analogy, triple bonds are stronger and shorter than doubles, because they contain one *σ* and two *π* bonds.

Example Question

Consider glycine, the simplest amino acid, whose structure is shown here:

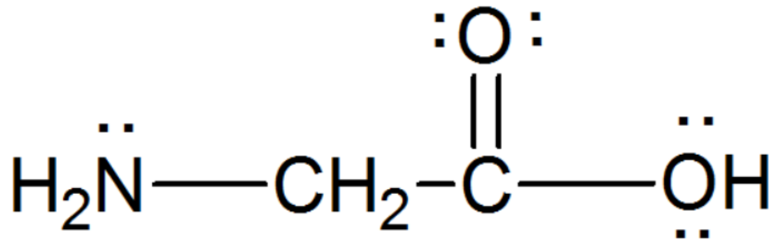


A. How many total σ bonds are there in glycine? **9**

B. How many total π bonds are there in glycine? **1**

Example Question

Consider glycine, the simplest amino acid, whose structure is shown here:



A. How many total σ bonds are there in glycine? **9**

B. How many total π bonds are there in glycine? **1**

Formal Charge

For some molecules, we can draw more than one ***Lewis structure*** that gives all the atoms a full octet. When this happens, we can think of the actual molecule being a “blend” of all of these “contributing” Lewis structures.

For example, we can draw two structures for CO_2 that both have a full octet on every atom:

So, which one is right?

Both of them. In reality, CO_2’s structure is somewhere in-between these two. We say, then, that these two structures are “contributors” (or ***resonance contributors***) to the ***actual*** structure of CO_2. However, the more stable of these two structures is the one that “contributes more.”

Formal Charge

So how do we know which one “contributes more?”

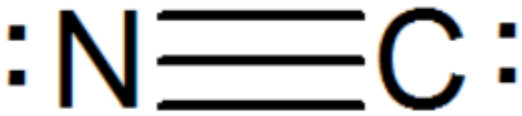
We assign numbers, called ***formal charges***, to each atom in the molecule. The Lewis structure that “contributes more” (is the most stable) is generally the one in which the atoms all have ***formal charges*** that are as close as possible to zero.

Sometimes, none of the ***contributors*** have a ***formal charge*** of zero on every single atom. When this happens, the most stable one is the one that gives negative ***formal charges*** to the most electronegative atoms and positive ***formal charges*** to the least electronegative atoms.

How to Determine Formal Charge

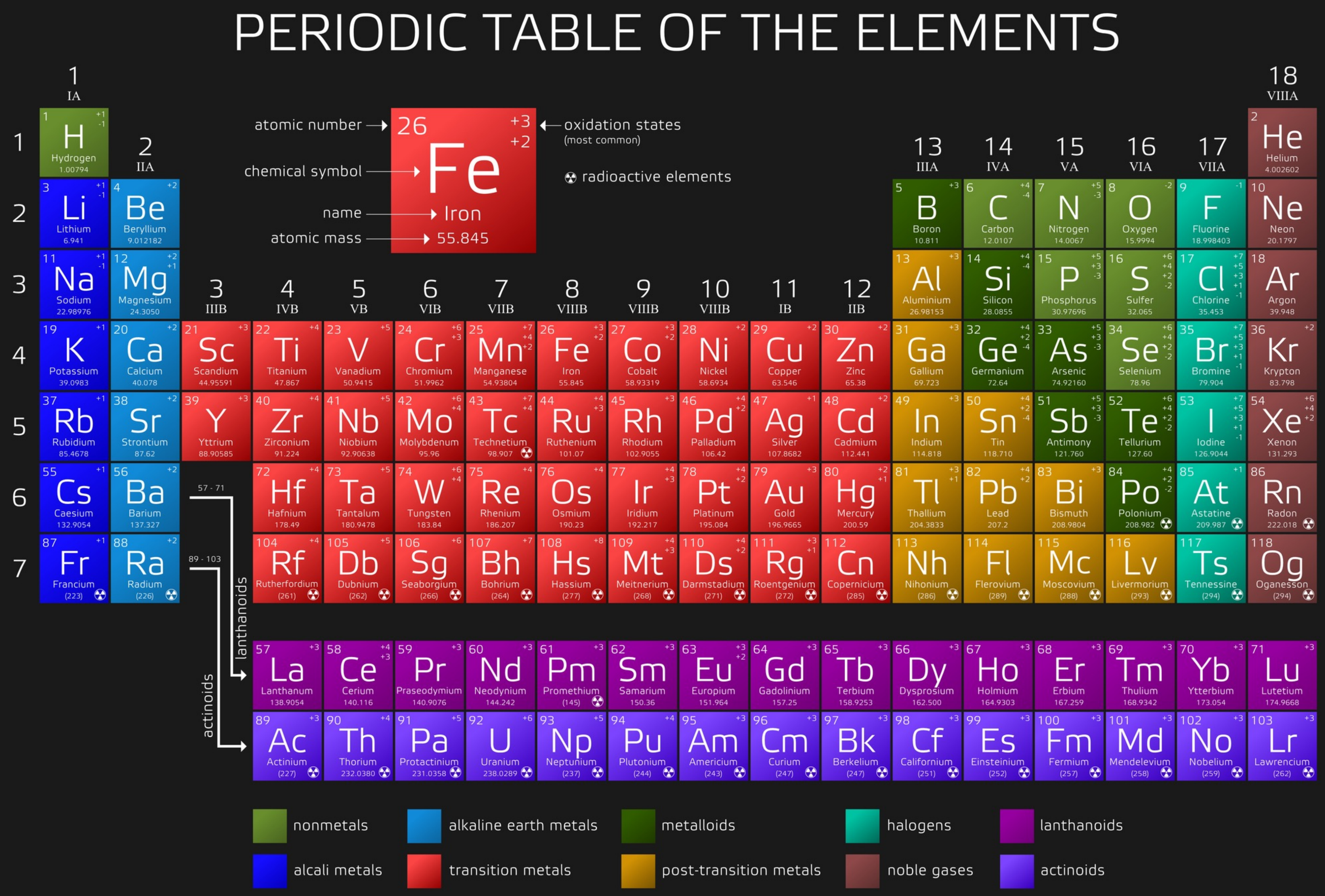
Formal Charge = (valence electrons) – (one for each dot) – (one for each bond)

What is the formal charge of the C and the N atom in cyanide?



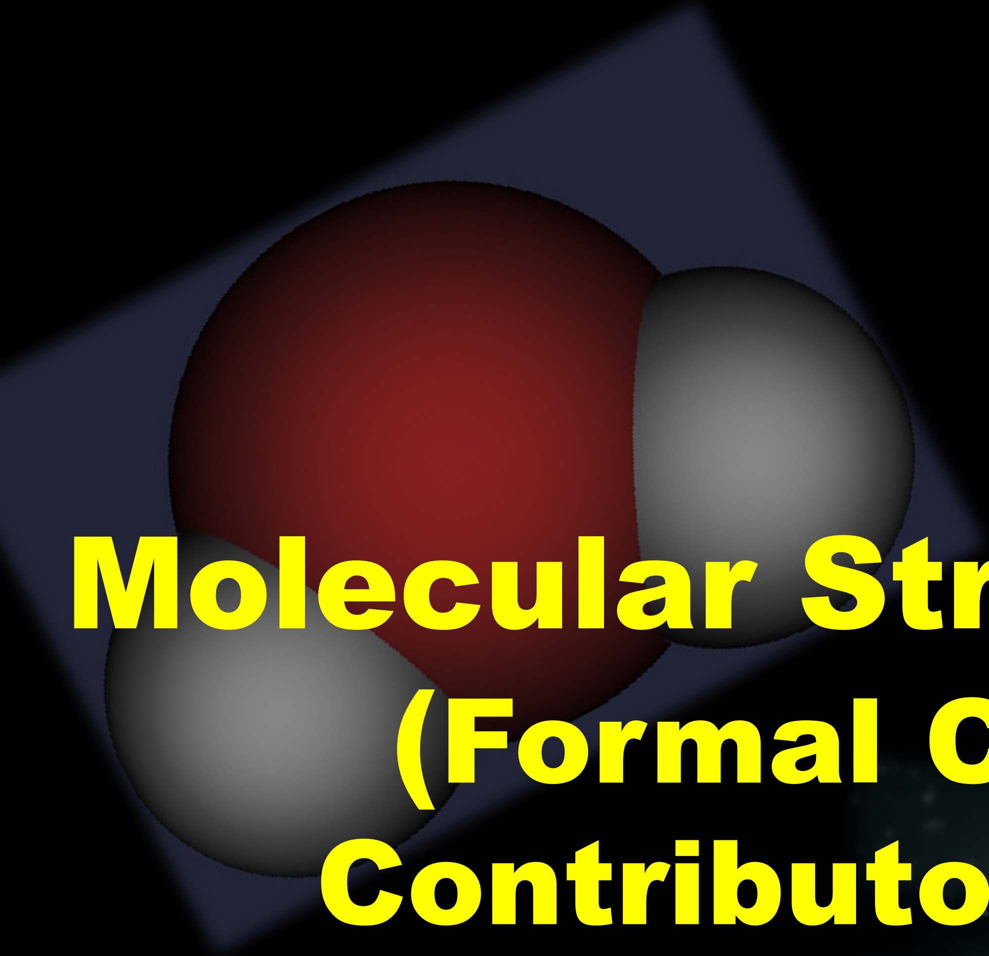
For **C**: **4 – (2) – (3) = –1**

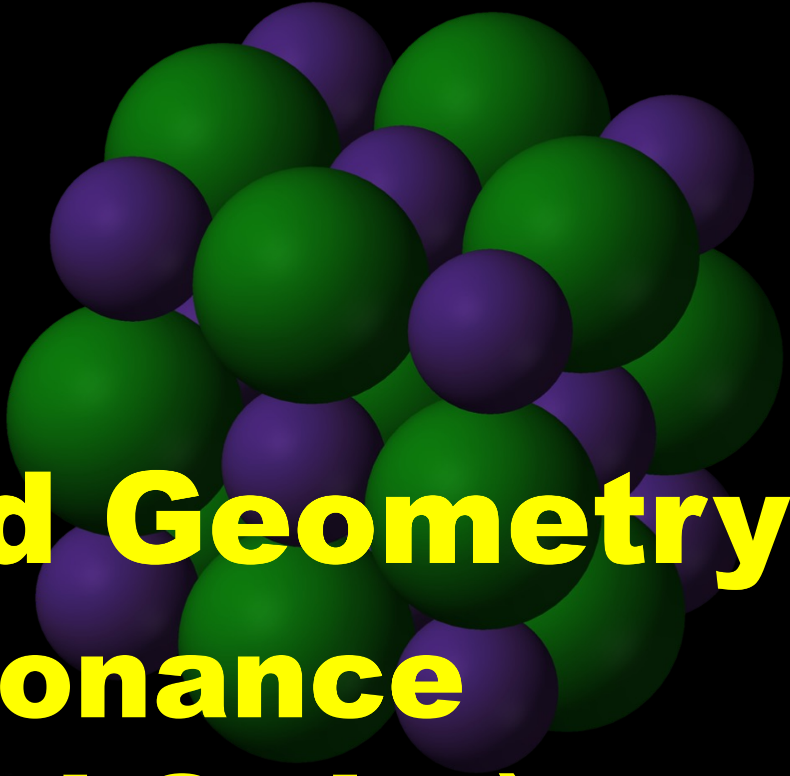
For **N**: **5 – (2) – (3) = 0**


PERIODIC TABLE OF THE ELEMENTS
atomic number
chemical symbol
name
atomic mass
oxidation states (most common)
26 Fe Iron 55.845 +3 +2
radioactive elements
1 IA
2 IIA
3 IIIB
4 IVB
5 VB
6 VIB
7 VIIB
8 VIIIB
9 VIIIB
10 VIIIB
11 IB
12 IIB
13 IIIA
14 IVA
15 VA
16 VIA
17 VIIA
18 VIIIA
1
2
3
4
5
6
7
1 H Hydrogen 1.00794 +1 -1
2 He Helium 4.002602
3 Li Lithium 6.941 +1 -1
4 Be Beryllium 9.012182 +2
5 B Boron 10.811 +3
6 C Carbon 12.0107 +4 -4
7 N Nitrogen 14.0067 +5 -3
8 O Oxygen 15.9994 -2
9 F Fluorine 18.998403 -1
10 Ne Neon 20.1797
11 Na Sodium 22.98976 +1 -1
12 Mg Magnesium 24.3050 +2 +1
13 Al Aluminium 26.98153 +3
14 Si Silicon 28.0855 +4 -4
15 P Phosphorus 30.97696 +5 +3 -3
16 S Sulfer 32.065 +6 +4 +2 -2
17 Cl Chlorine 35.453 +7 +5 +3 +1 -1
18 Ar Argon 39.948
19 K Potassium 39.0983 +1
20 Ca Calcium 40.078 +2
21 Sc Scandium 44.95591 +3
22 Ti Titanium 47.867 +4
23 V Vanadium 50.9415 +5
24 Cr Chromium 51.9962 +6 +3
25 Mn Manganese 54.93804 +7 +4 +2
26 Fe Iron 55.845 +3 +2
27 Co Cobalt 58.93319 +3 +2
28 Ni Nickel 58.6934 +2
29 Cu Copper 63.546 +2
30 Zn Zinc 65.38 +2
31 Ga Gallium 69.723 +3
32 Ge Germanium 72.64 +4 +2 -4
33 As Arsenic 74.92160 +5 +3 -3
34 Se Selenium 78.96 +6 +4 +2 -2
35 Br Bromine 79.904 +7 +5 +3 +1 -1
36 Kr Krypton 83.798 +2
37 Rb Rubidium 85.4678 +1
38 Sr Strontium 87.62 +2
39 Y Yttrium 88.90585 +3
40 Zr Zirconium 91.224 +4
41 Nb Niobium 92.90638 +5
42 Mo Molybdenum 95.96 +6 +4
43 Tc Technetium 98.907 +7 +4
44 Ru Ruthenium 101.07 +4 +3
45 Rh Rhodium 102.9055 +3
46 Pd Palladium 106.42 +4 +2
47 Ag Silver 107.8682 +1
48 Cd Cadmium 112.441 +2
49 In Indium 114.818 +3
50 Sn Tin 118.710 +4 +2 -4
51 Sb Antimony 121.760 +5 +3 -3
52 Te Tellurium 127.60 +6 +4 +2 -2
53 I Iodine 126.9044 +7 +5 +3 +1 -1
54 Xe Xenon 131.293 +6 +4 +2
55 Cs Caesium 132.9054 +1
56 Ba Barium 137.327 +2
57 - 71
72 Hf Hafnium 178.49 +4
73 Ta Tantalum 180.9478 +5
74 W Tungsten 183.84 +6 +4
75 Re Rhenium 186.207 +4
76 Os Osmium 190.23 +4
77 Ir Iridium 192.217 +4 +3
78 Pt Platinum 195.084 +4 +2
79 Au Gold 196.9665 +3
80 Hg Mercury 200.59 +2 +1
81 Tl Thallium 204.3833 +3 +1
82 Pb Lead 207.2 +4 +2
83 Bi Bismuth 208.9804 +3
84 Po Polonium 208.982 +4 +2 -2
85 At Astatine 209.987 +1
86 Rn Radon 222.018
87 Fr Francium (223) +1
88 Ra Radium (226) +2
89 - 103
104 Rf Rutherfordium (261) +4
105 Db Dubnium (262) +5
106 Sg Seaborgium (266) +6
107 Bh Bohrium (264) +7
108 Hs Hassium (277) +8
109 Mt Meitnerium (268) +4 +3
110 Ds Darmstadium (271) +4 +2
111 Rg Roentgenium (272) +3 +1
112 Cn Copernicium (285) +2
113 Nh Nihonium (286)
114 Fl Flerovium (289)
115 Mc Moscovium (288)
116 Lv Livermorium (293)
117 Ts Tennessine (294)
118 Og Oganesson (294)
lanthanoids
57 La Lanthanum 138.9054 +3
58 Ce Cerium 140.116 +4 +3
59 Pr Praseodymium 140.9076 +3
60 Nd Neodymium 144.242 +3
61 Pm Promethium (145) +3
62 Sm Samarium 150.36 +3
63 Eu Europium 151.964 +3 +2
64 Gd Gadolinium 157.25 +3
65 Tb Terbium 158.9253 +3
66 Dy Dysprosium 162.500 +3
67 Ho Holmium 164.9303 +3
68 Er Erbium 167.259 +3
69 Tm Thulium 168.9342 +3
70 Yb Ytterbium 173.054 +3
71 Lu Lutetium 174.9668 +3
actinoids
89 Ac Actinium (227) +3
90 Th Thorium 232.0380 +4
91 Pa Protactinium 231.0358 +5
92 U Uranium 238.0289 +6
93 Np Neptunium (237) +5
94 Pu Plutonium (244) +4
95 Am Americium (243) +3
96 Cm Curium (247) +3
97 Bk Berkelium (247) +3
98 Cf Californium (251) +3
99 Es Einsteinium (252) +3
100 Fm Fermium (257) +3
101 Md Mendelevium (258) +3
102 No Nobelium (259) +3
103 Lr Lawrencium (262) +3
nonmetals
alkaline earth metals
metalloids
halogens
lanthanoids
alcali metals
transition metals
post-transition metals
noble gases
actinoids

Example Question

What is the formal charge of each C and O in the following ***Lewis structures?*** Which ***Lewis structure*** is the greater resonance contributor?

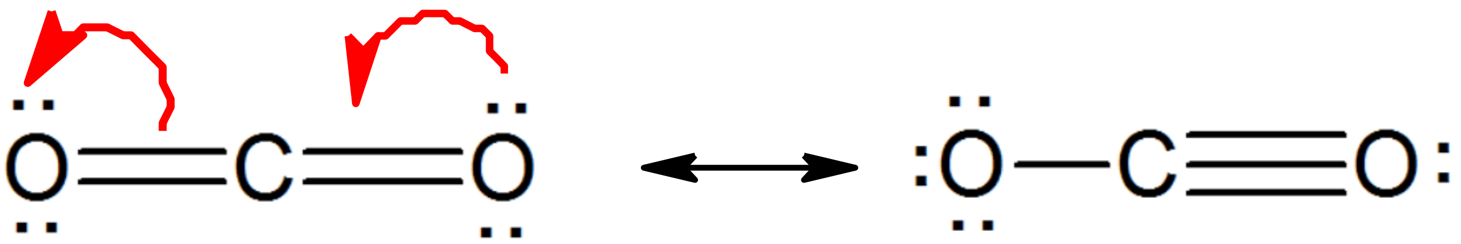




Molecular Structure and Geometry (Formal Charge, Resonance Contributors, and Bond Order)

Bond Order

To be honest, ***bond order*** is a subject that can get a little complicated. To keep it simple, just think of a bond's ***bond order*** as being its average length, when multiple ***contributors*** are possible. The larger the ***bond order***, the stronger the bond. For example, CO_2 has three major ***resonance contributors:***



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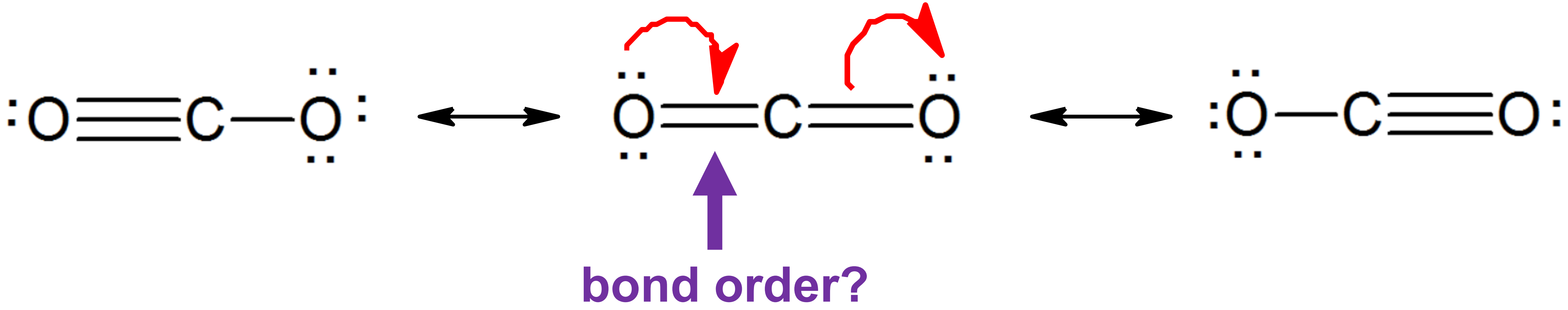

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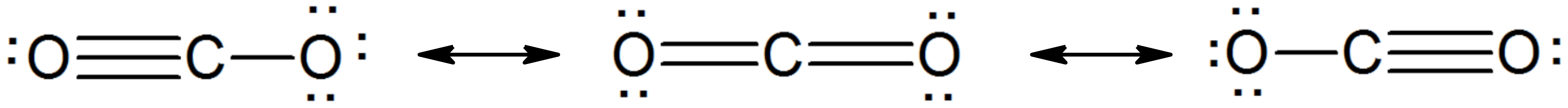


Bond order **= average of 2, 3, and 1**

= (2 + 3 + 1) / 3 = 2

Example Question

The structure of the carbonate ion, CO_3^{2-}, can be represented by the ***Lewis structure*** shown below. What is the C–O bond order in CO_3^{2-}?

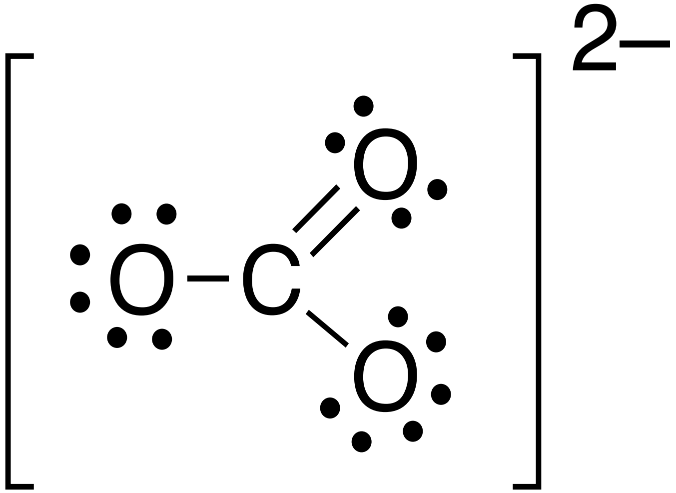
A. 1

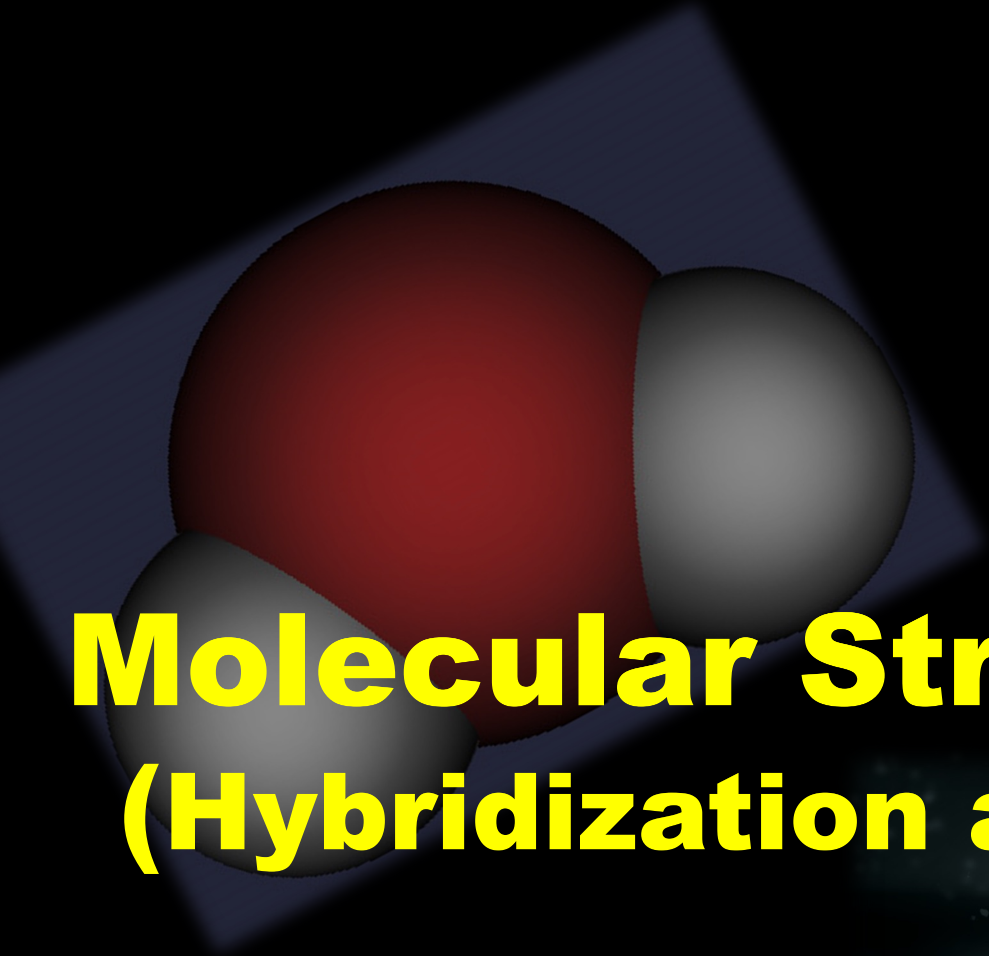
B. 1.3

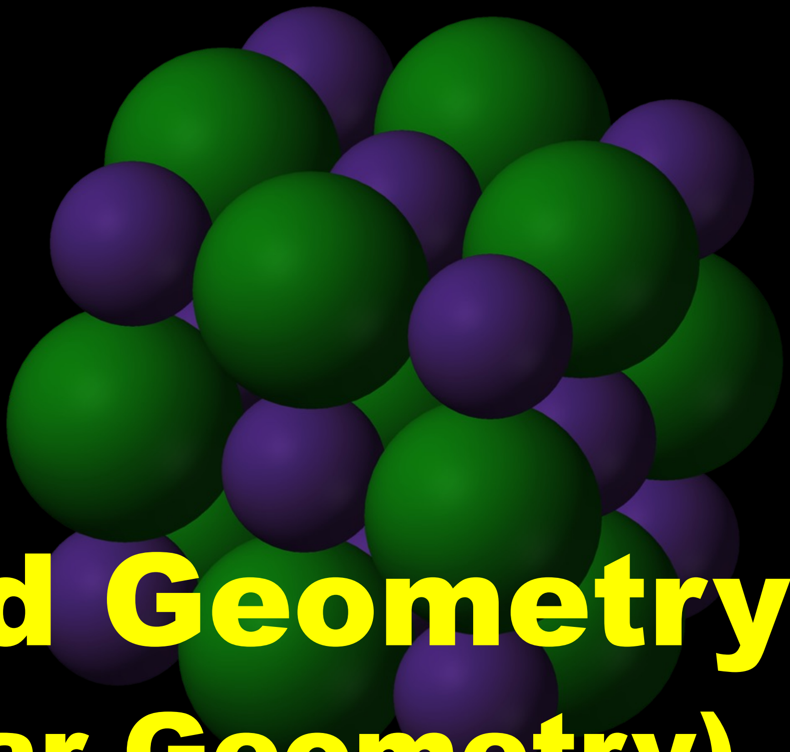
C. 1.5

D. 2

E. 2.5





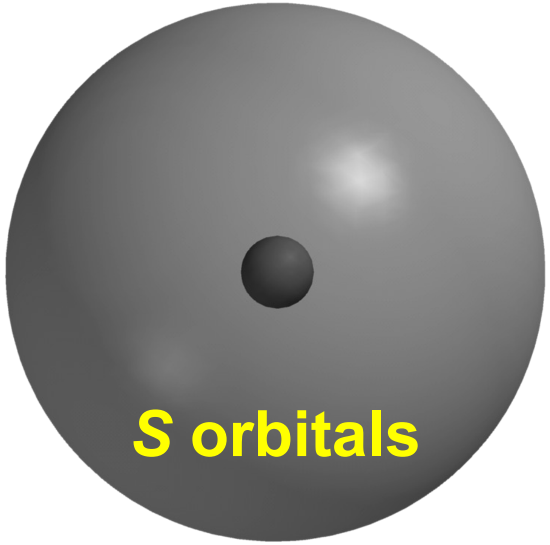


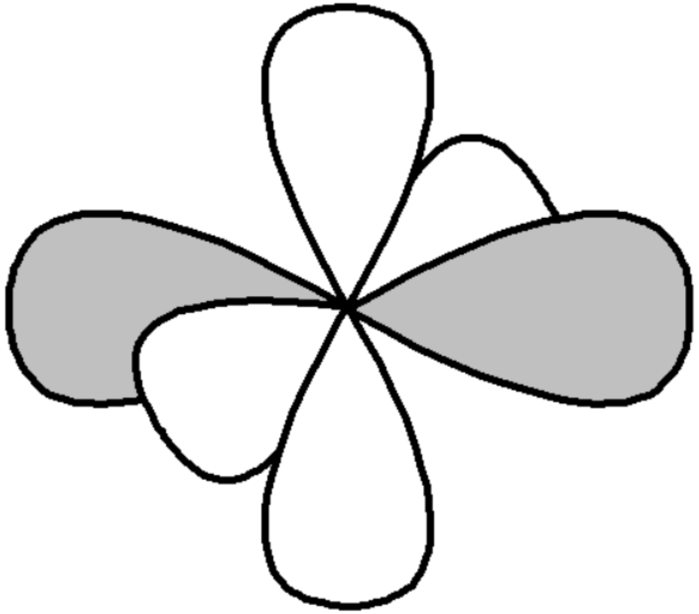
Molecular Structure and Geometry

(Hybridization and Molecular Geometry)

Orbital Recap

In an earlier chapter, I taught you that atoms place their various electrons in . . .

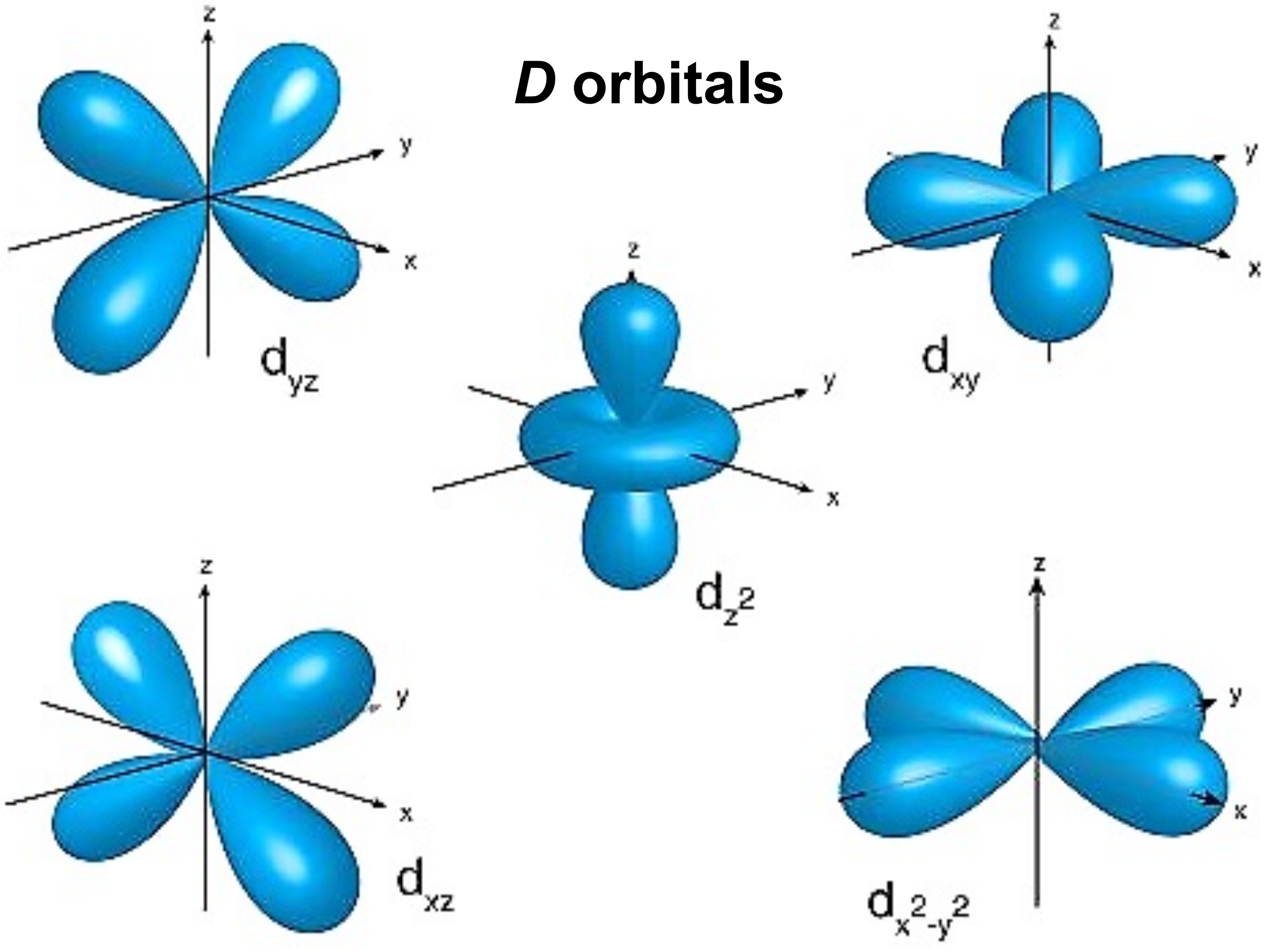




***P* orbitals**

Orbital Recap

In an earlier chapter, I taught you that atoms place their various electrons in . . .

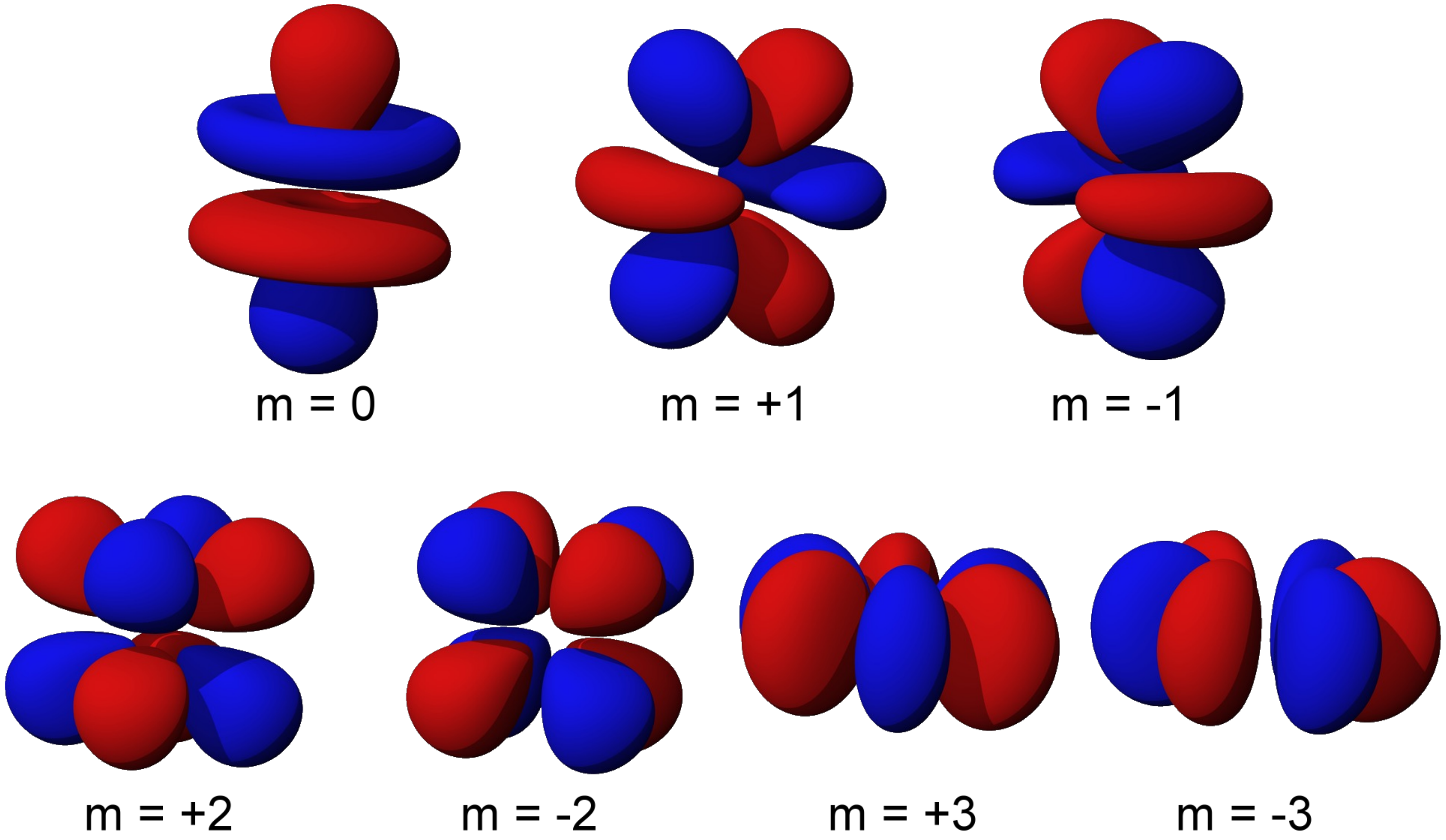


***D* orbitals**

Orbital Recap

In an earlier chapter, I taught you that atoms place their various electrons in . . .

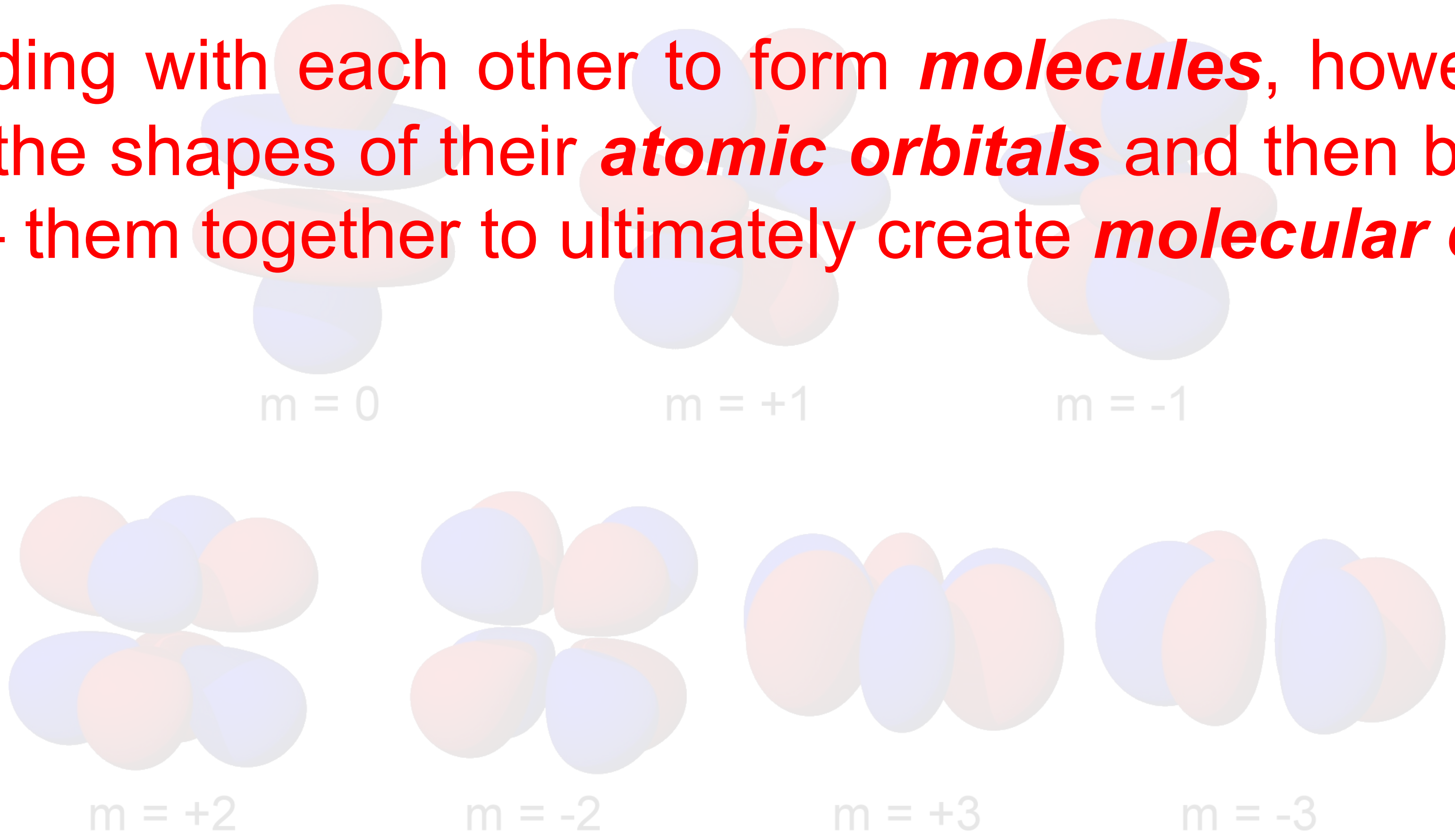
F orbitals



Orbital Recap

In an earlier chapter, I taught you that atoms place their various electrons in . . .

***F* orbitals**

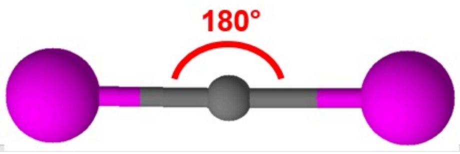


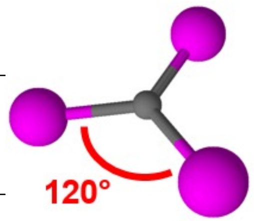
When bonding with each other to form ***molecules***, however, atoms have to change the shapes of their ***atomic orbitals*** and then blend and mix –or ***hybridize***– them together to ultimately create ***molecular orbitals***.

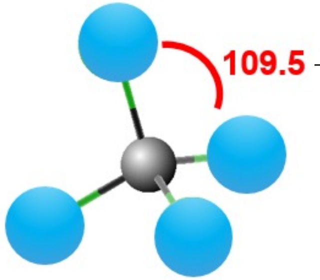
Orbital Hybridization

To identify any atom's hybridization, you just have to count the number of "things" around the atom* and then memorize the following:

# of things	Hybridization
2	sp
3	sp^2
4	sp^3







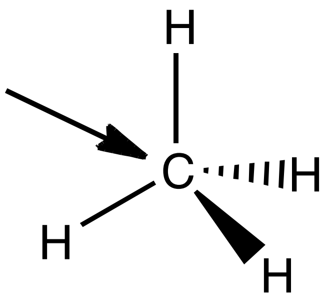
*A "thing" here is either another atom or a lone electron pair (1 nonbonding pair = 1 thing).

Orbital Hybridization

What is the hybridization of each of the indicated atoms?

hybridization?

***sp*^{3}**

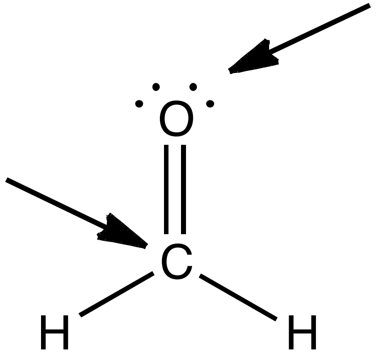


hybridization?

***sp*^{2}**

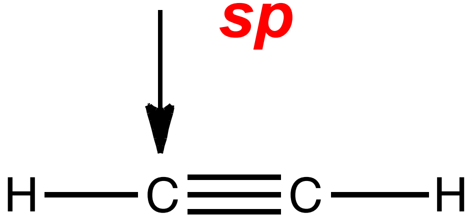
hybridization?

***sp*^{2}**

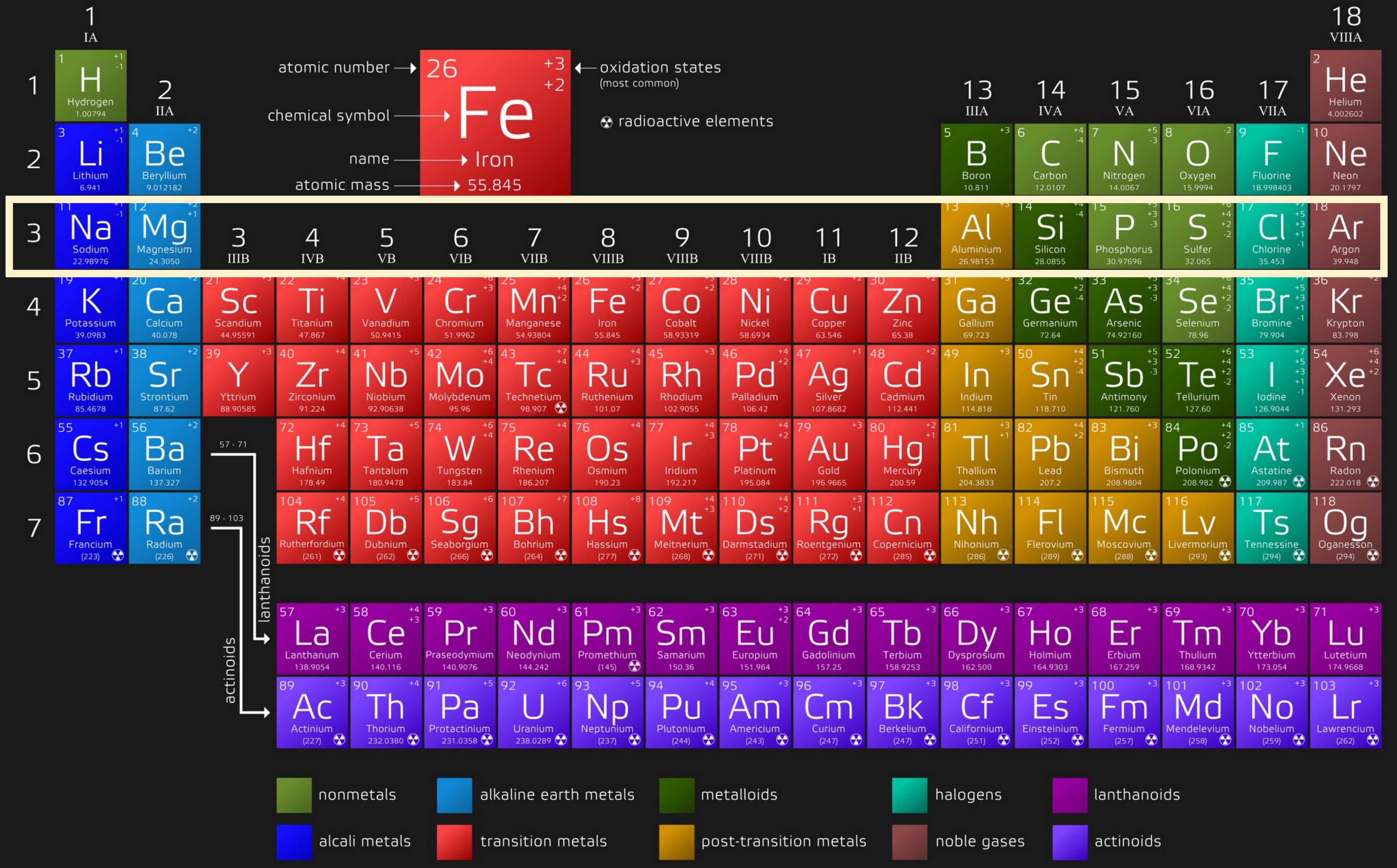


hybridization?

sp



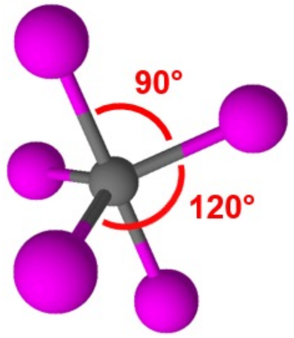
PERIODIC TABLE OF THE ELEMENTS

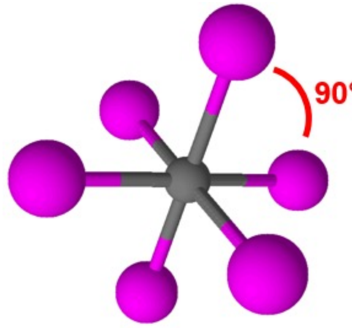


More Complex Hybridization

To identify any atom's hybridization, you just have to count the number of "things" around the atom* and then memorize the following:

# of things	Hybridization
5	sp^3d
6	sp^3d^2

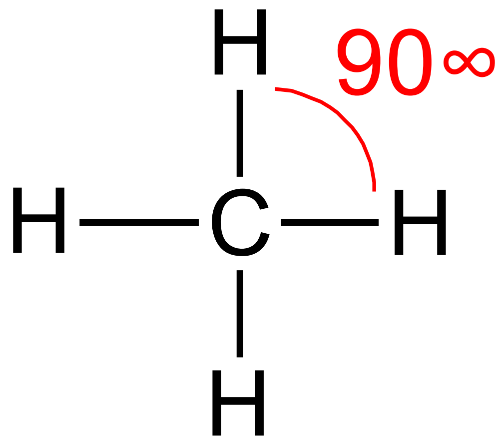




*A "thing" here is either another atom or a lone electron pair (1 nonbonding pair = 1 thing).

Molecular Geometry (Shape)

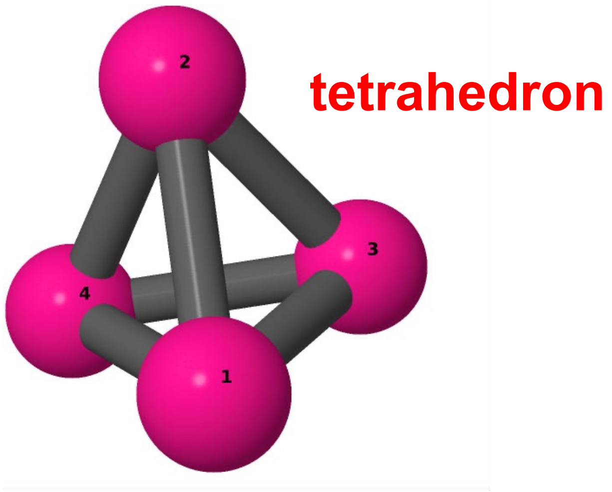
Generally speaking, molecules adopt the shape that keeps all the atoms as far away (as spaced out) from each other as possible. For example, in the molecule CH_4 . . .

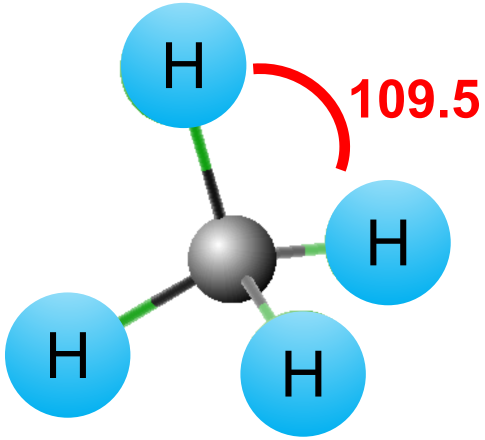


. . . The furthest apart that the four hydrogen atoms can get on a flat, two-dimensional surface (like this lecture slide) is 90 degrees. The idea that atoms spread out as much as possible is called the ***VSEPR*** (Valence-Shell Electron-Pair Repulsion) ***model***.

Molecular Geometry (Shape)

In real life, however, molecules do not exist in a two-dimensional world. They exist in a three-dimensional world, where each of these hydrogen atoms can actually spread out further than 90°, like this:

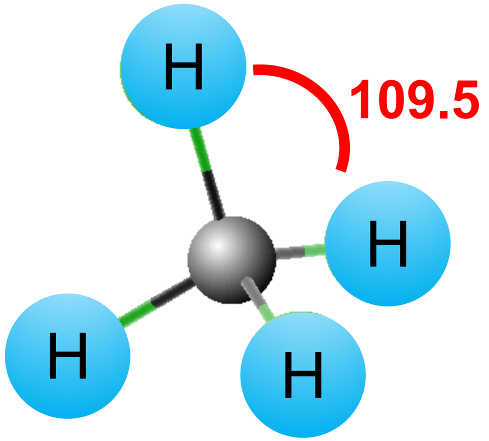


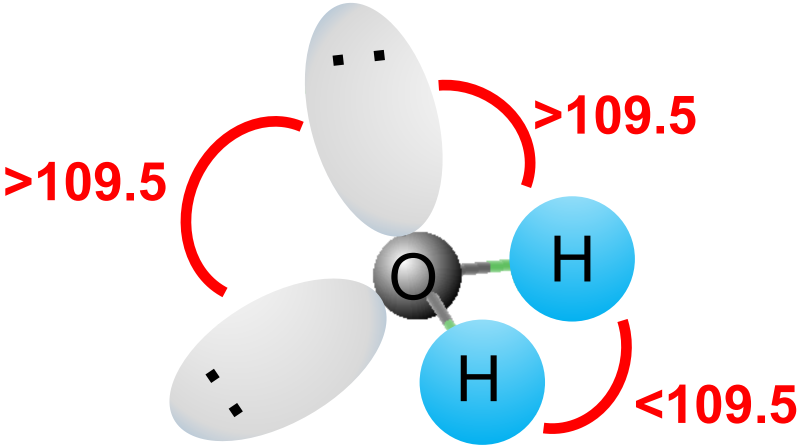


This type of shape is called a ***tetrahedron***. You can see that the hydrogens (blue balls) can spread out from their central carbon (the gray ball) to have a 109.5° angle between each hydrogen atom.

H_2O vs. CH_4

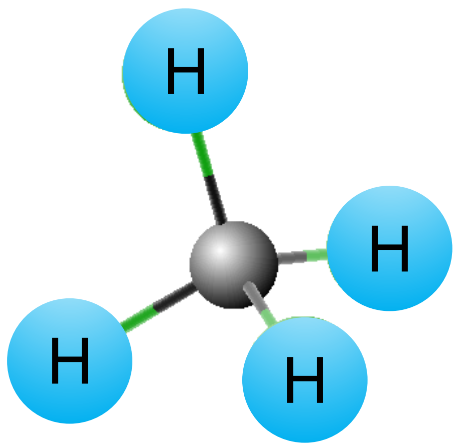
Like CH_4, H_2O also has four things around its central oxygen atom. However, two of those things are lone pairs:



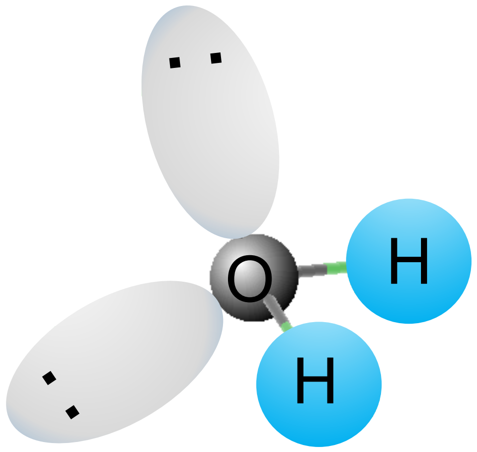


Lone pairs take up more room than regular bonding groups. Thus, in H_2O, te angles between the hydrogen atoms is <u><109.5</u>, while remaining angles are <u>>109.5</u>.

What is H_2O's Shape?



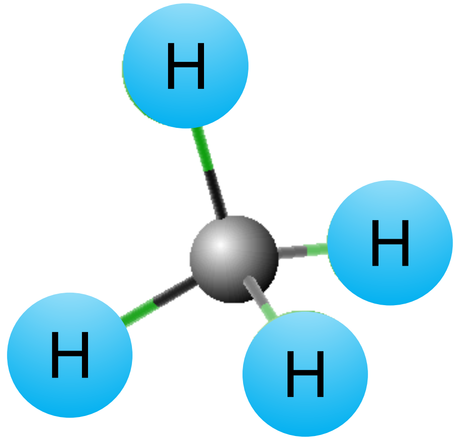
tetrahedron



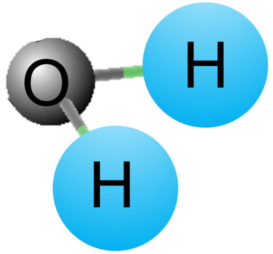
e^- included: tetrahedron

What is H_2O's Shape?

e^- included: ***electron-domain geometry***
e^- NOT included: ***molecular geometry***

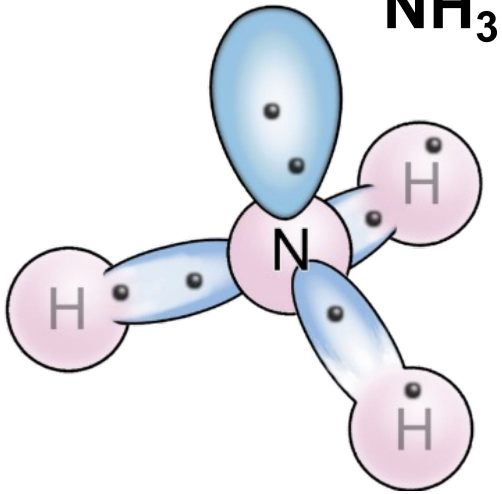


tetrahedron



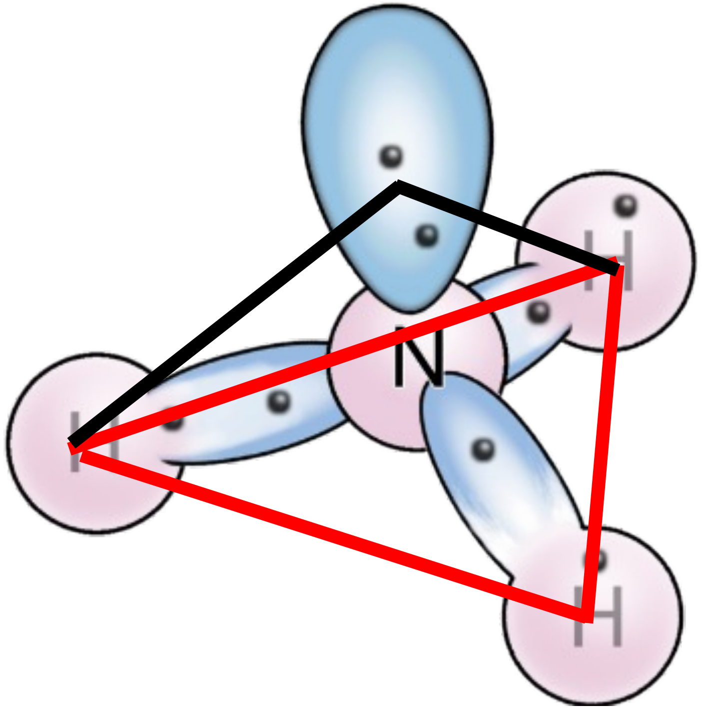
e^- included: tetrahedron
e^- NOT included: bent

NH_3



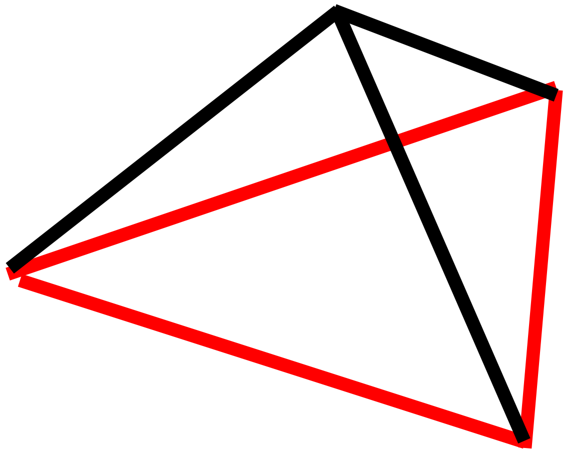
e^- included: tetrahedron
e^- NOT included: ???

Ammonia (NH_3)



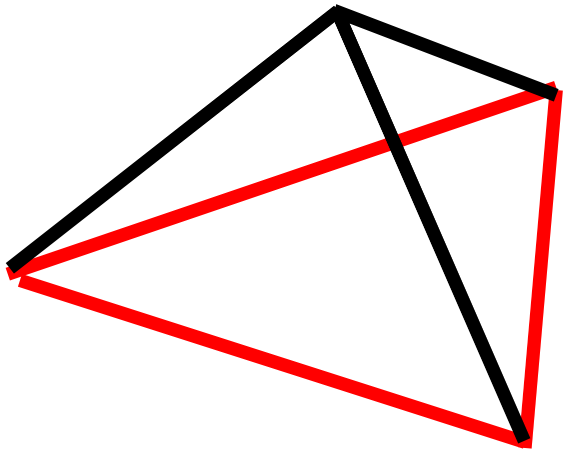
e^- included (electron-domain geometry): tetrahedron
e^- NOT included (molecular geometry): ???

Ammonia (NH_3)



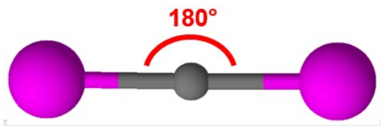
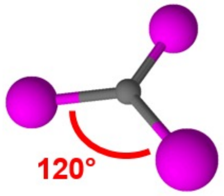
e^- included (electron-domain geometry): tetrahedron
e^- NOT included (molecular geometry): ???

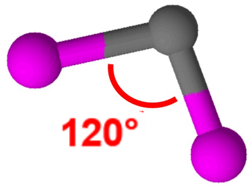
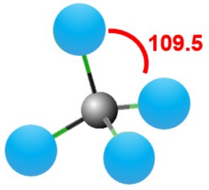
Ammonia (NH_3)

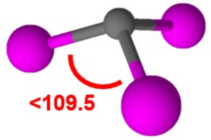


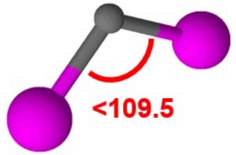
e⁻ included (electron-domain geometry): tetrahedron
e⁻ NOT included (molecular geometry): trigonal pyramid

Memorize This Information

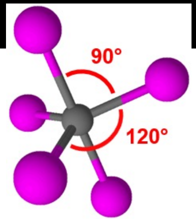
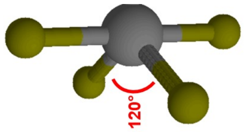
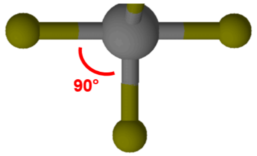
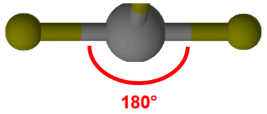
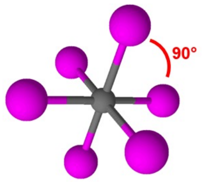
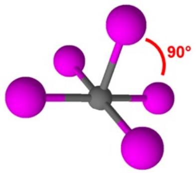
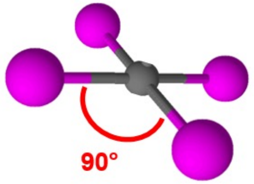
<table>
<tr><th>Electron Domain (‘things’)</th><th>Hybrid-ization</th><th>Bond Angles</th><th>Electron-Domain Geometry</th><th>Non-Bonding Electron Pairs</th><th>Shapes</th><th>Molecular Geometry</th></tr>
<tr><td>2</td><td>sp</td><td>180°</td><td>Linear</td><td>0</td><td>
</td><td>Linear</td></tr>
<tr><td rowspan="2">3</td><td rowspan="2">sp^2</td><td rowspan="2">120°</td><td rowspan="2">Trigonal planar</td><td>0</td><td rowspan="2">


</td><td>Trigonal planar</td></tr>
<tr><td>1</td><td>bent</td></tr>
<tr><td rowspan="3">4</td><td rowspan="3">sp^3</td><td rowspan="3">109.5°</td><td rowspan="3">Tetra-hedral</td><td>0</td><td rowspan="3">




</td><td>Tetrahedral</td></tr>
<tr><td>1</td><td>Trigonal pyramid</td></tr>
<tr><td>2</td><td>bent</td></tr>
</table>

Memorize This Information

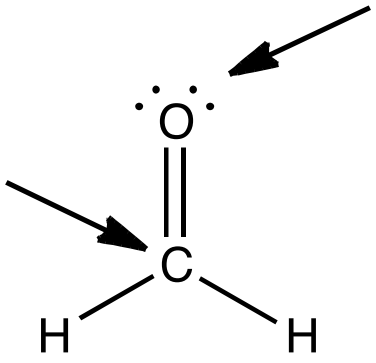
Electron Domain ('things')	Hybrid-ization	Bond Angles	Electron-Domain Geometry	Non-Bonding Electron Pairs	Shape	Molecular Geometry
5	***sp^3d***	90° 120° 180°	Trigonal bipyramid	0		Trigonal bipyramid
				1		See-saw
				2		T-shaped
				3		Linear
6	***sp^3d^2***	90°	Octa-hedral	0		Octahedral
				1		Square pyramid
				2		Square planar

Molecular Geometry (Shape)

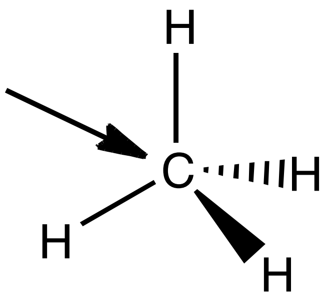
What is the bond angle and geometry of the indicated atoms?

~120°
Trigonal plane

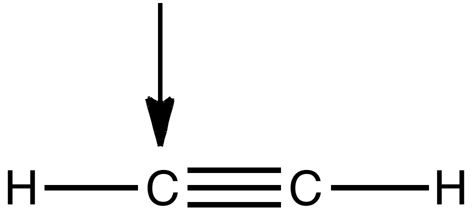
120°
Trigonal plane

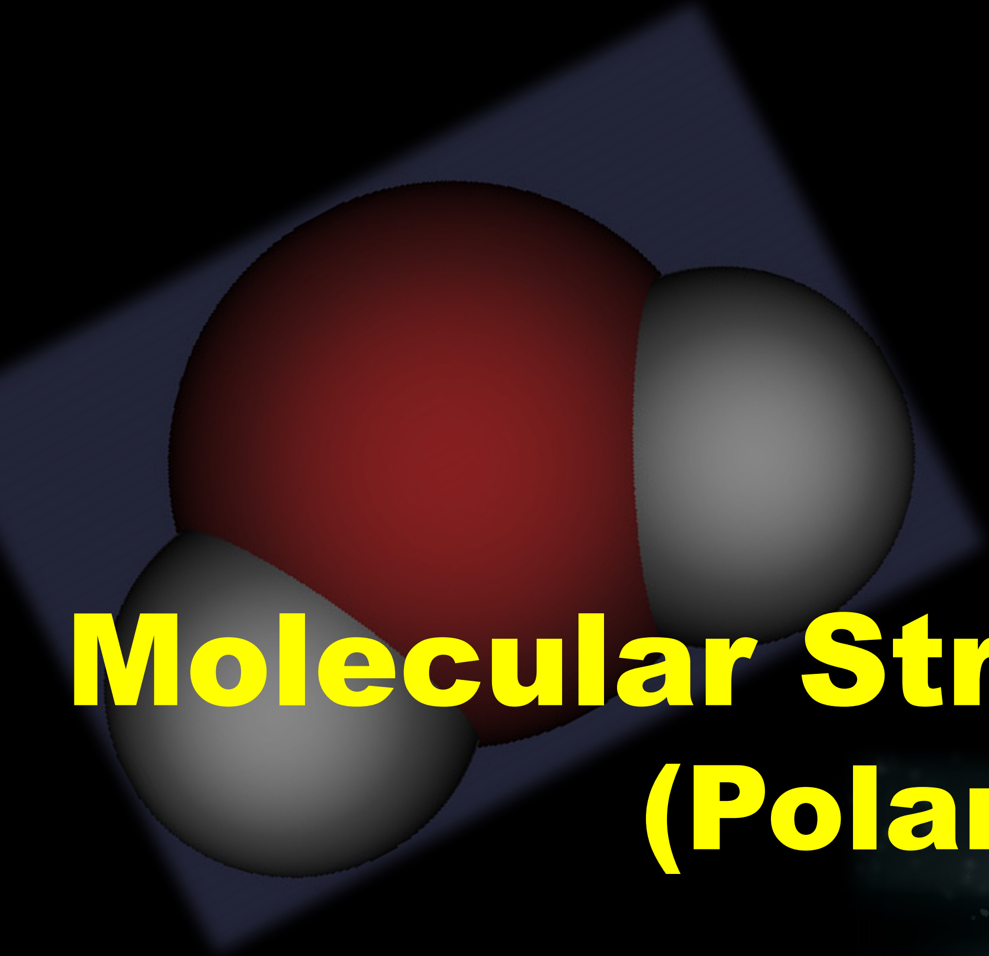


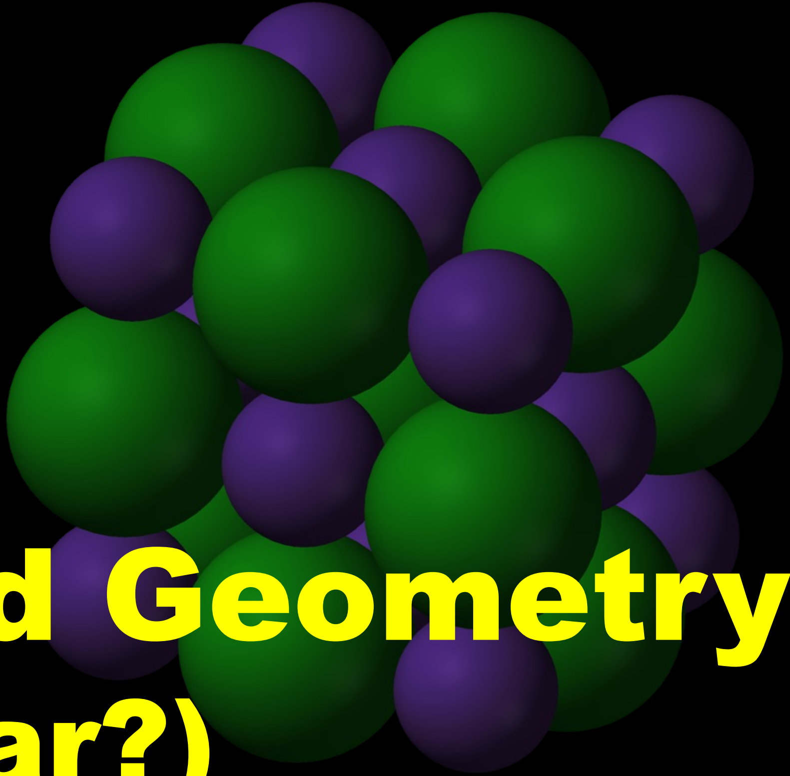
109.5°
tetrahedron



180°
linear





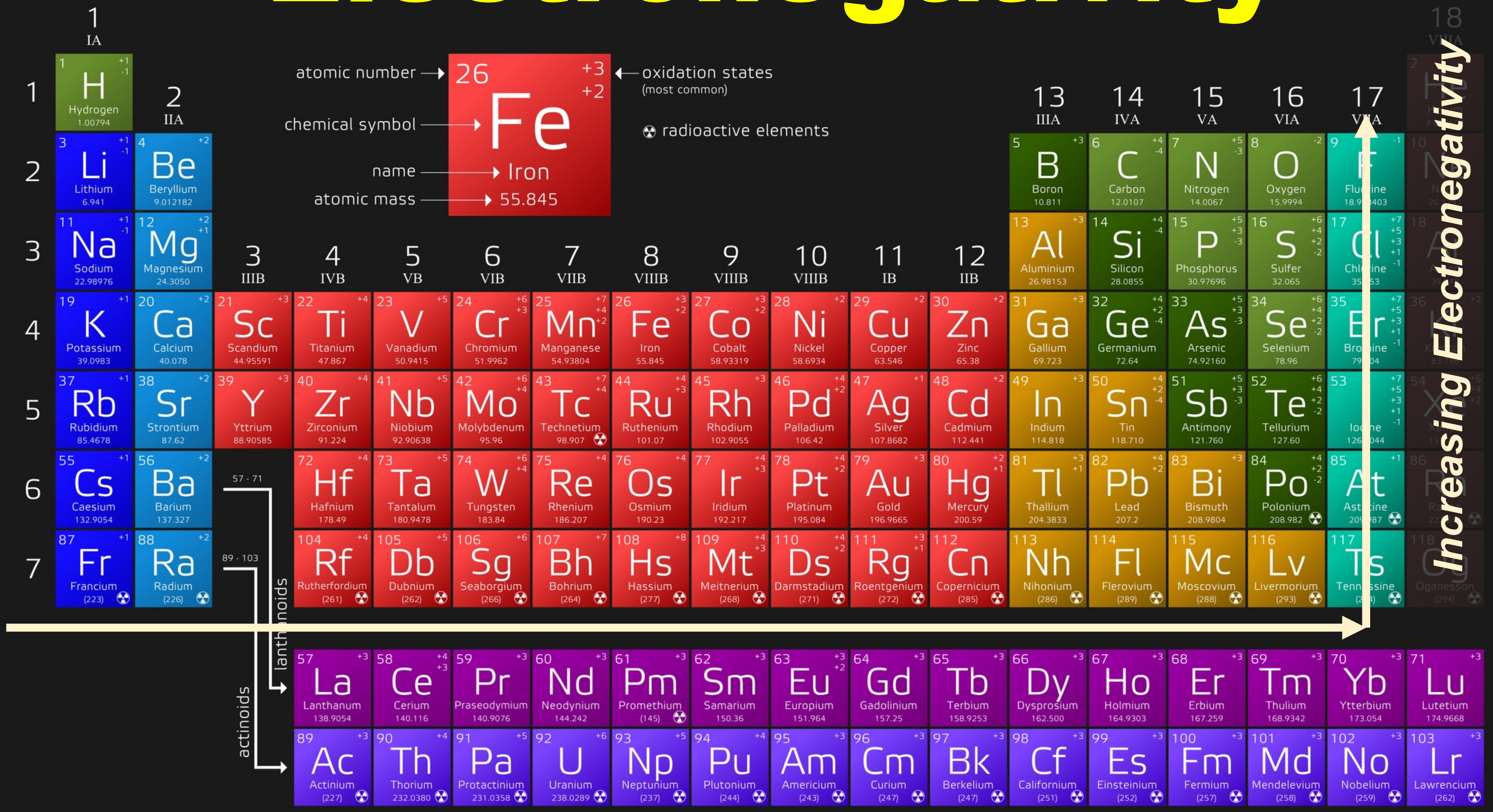


Molecular Structure and Geometry (Polar or Nonpolar?)

Polar or Nonpolar?

As I discussed in an earlier video, when bonded atoms have a significant difference in their electronegativities, they share their electrons unequally. This causes their bond to be polar. We should, of course, remember that . . .

Electronegativity

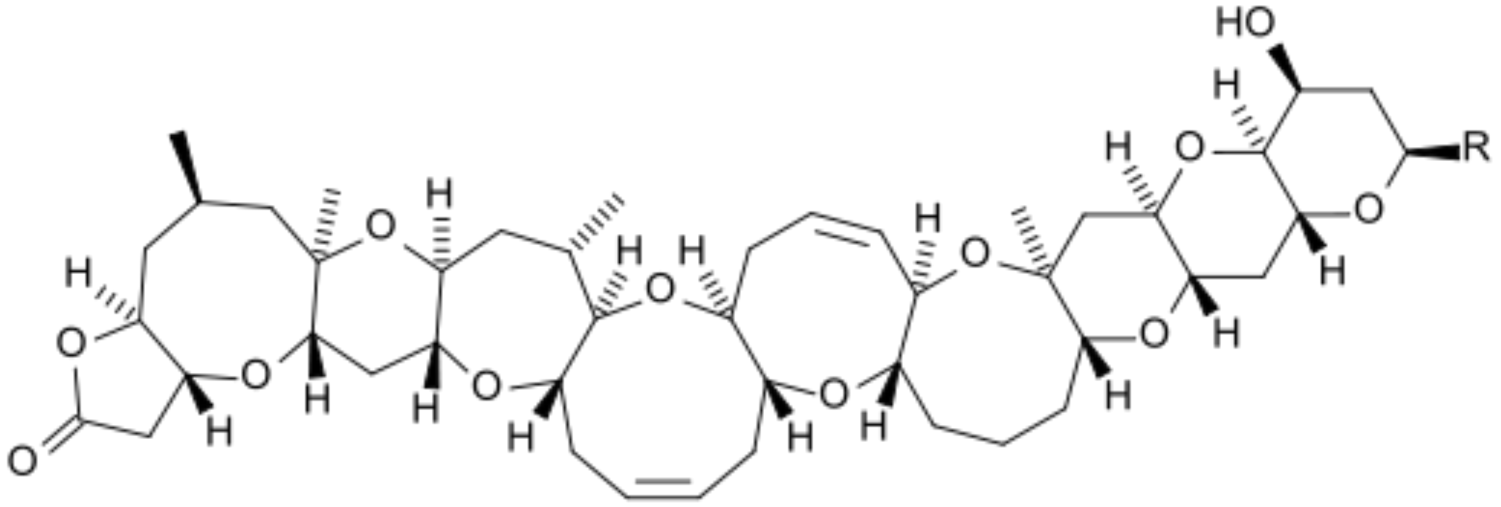


. . . An atom's ***electronegativity*** (its "thirst" for electrons) increases as you go up and to the right on the periodic table.

Polar or Nonpolar?

As I discussed in an earlier video, when bonded atoms have a significant difference in their electronegativities, they share their electrons unequally. This causes their bond to be polar. We should, of course, remember that . . .

Determining the ***polarity*** of entire complex molecules is complicated, as their ***polarities*** are the collective product of the combined polarities and geometries of their individual bonds.



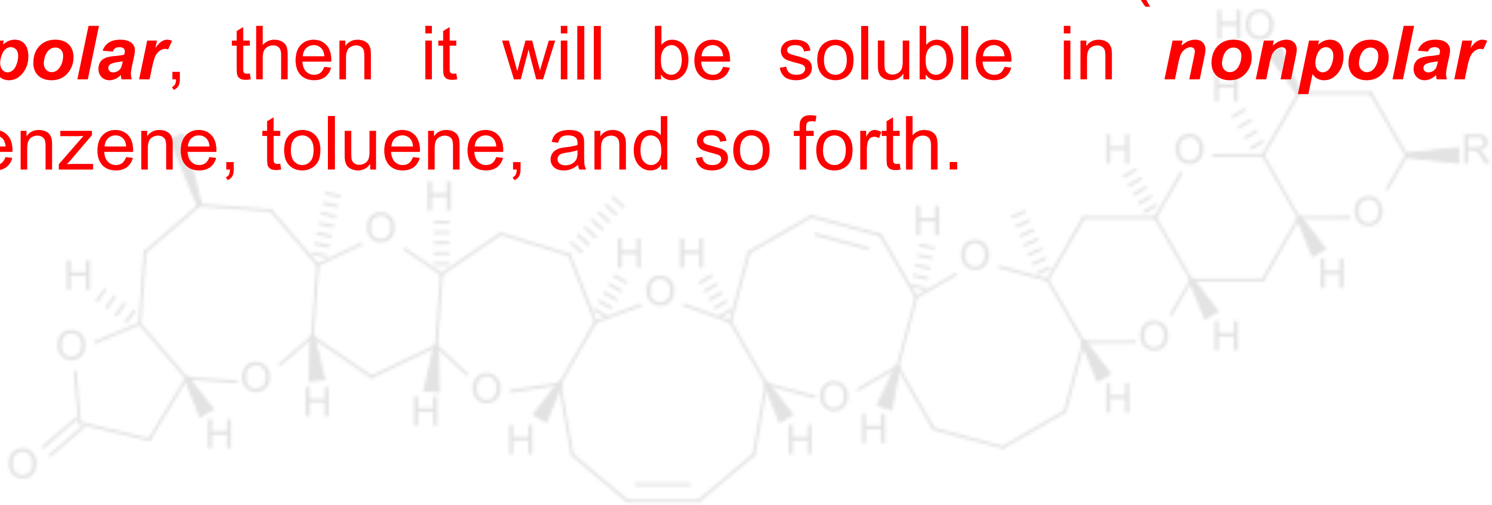
Brevetoxin A

Polar or Nonpolar?

As I discussed in an earlier video, when bonded atoms have a significant difference in their electronegativities, they share their electrons unequally. This causes their bond to be polar. We should, of course, remember that ...

We care about ***polarity*** because ***polarity*** determines solubility. In other words, if something is ***polar*** then it will be soluble in water (and other polar solvents). If it's ***nonpolar***, then it will be soluble in ***nonpolar*** solvents, such as hexane, benzene, toluene, and so forth.

Determining the ***polarity*** of entire complex molecules is complicated, as their ***polarities*** are the collective product of the combined polarities and geometries of their individual bonds.



Brevetoxin A

Polar or Nonpolar?

For simpler molecules, we can more easily determine if they are ***polar*** or ***nonpolar***. Here are some tips:

- O–H bonds are polar. The more OH's your molecule has, the more polar it is.

- C–H bonds are nonpolar. The more CH's your molecule has, the more nonpolar it is.

For other simple molecules, I'll now give you some quick rules.

Polar or Nonpolar?

So how do we determine if a molecule is polar or not?

1. Draw your molecule's ***Lewis structure*** flat on your paper and then spread out all of its groups (including its lone pairs) as far as possible, around its central atom.

2. Draw arrows between every atom in the molecule, going from the less electronegative atom (**A**) to the more electronegative atom (**B**) in each bond, like this:

A – B

3. Answer the following question: "If my central atom were a truck stuck in the mud being pulled in the directions indicated by the arrows, would the truck move?" so, then your molecule is ***polar***. If not, then it's ***nonpolar***.

Example Question

Which of the following molecules would be most soluble in toluene? Which one would be the least?

a. $CH_3CH_2CH_2OH$

b. CH_3CH_2OH

c. $CH_3CH_2CH_2CH_2OH$

d. CH_3OH

Example Question

Predict whether each of the following molecules is ***polar*** or ***nonpolar***:

a. CCl_4

b. NH_3

c. SF_4

d. XeF_4