Name Completed Notes Hour

Unit 7: Lewis Structures, Shape, and Bond Strength

Learning Goals:

1. I can explain different characteristics of ionic and covalent substances.
2. I can calculate the number of valence electrons for an atom and a compound.
3. I can draw a Lewis structure for a molecule.
4. I can determine the molecular shape of a Lewis structure.
5. I can order bonds from strongest to weakest, including ionic bonds and every kind of covalent bond.

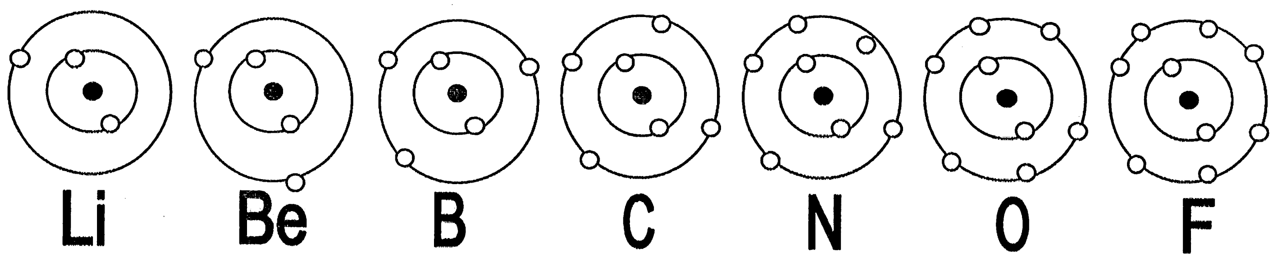
VOCABULARY (I can define/describe the following terms in my own words)

- bent shape
- bond length
- bonding site
- bond strength
- covalent bond
- double bond
- ionic bond
- linear shape
- lewis structures
- shared electron pairs
- single bond
- tetrahedral shape
- trigonal planar
- trigonal pyramidal
- triple bond
- unshared electron pairs (lone pairs)
- VSEPR

POGIL #1 – Lewis Dot Structures

Information: Bohr Diagrams

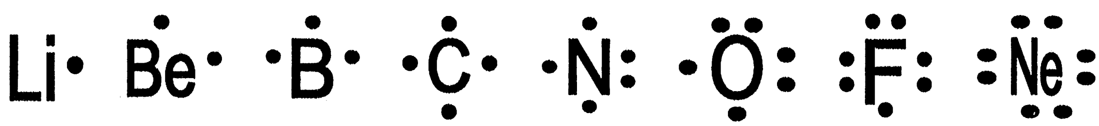
Below are seven "Bohr diagrams" for atoms #3-9.
FIGURE 1:



Information: Electron Dot Diagrams

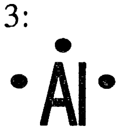
Below are electron dot diagrams, also known as "Lewis Structures" for atoms #3-10.

FIGURE 2:



The position of the dots is important. For example, another atom that has three dots in its Lewis structure is aluminum. Aluminum's three dots must be positioned the same way as boron's. Thus, aluminum's Lewis structure is:

FIGURE 3:



Critical Thinking Questions

8. What relationship exists between an atom's valence electrons and the number of dots in the Lewis structure of the atom?

   They are the same number

9. Why does nitrogen's Lewis Structure has five dots around it while nitrogen's Bohr diagram contains 7 dots around it.

   Because nitrogen has 5 valence electrons

10. Recall from questions 1-4 that oxygen, sulfur and selenium all have the same number of valence electrons (6). They also are in the same column of the periodic table. Predict how many valence electrons tellurium (Te) will have. 6

11. Comparing Figure 2 and Figure 3 we see that boron and aluminum have the same number of dots in their Lewis structures. Notice they are in the same column of the periodic table. Write the Lewis structure for gallium (Ga). ·Ga·

12. Write the Lewis structure for sulfur and selenium. Compare the structures you write with oxygen's Lewis structure from Figure 2.

a) Sulfur

S

b) Selenium

Se

13. In question seven, you drew the Bohr diagram for magnesium. Now write the Lewis Structure for magnesium. What similarities exist between the Lewis Structures for magnesium and beryllium?

Mg (Same number of dots)

Be (Same number of valence electrons)

14. Complete this statement: If elements are in the same column of the periodic table, they must have Lewis Structures that are Similar (similar or different).

15. Why does sodium have the same Lewis structure as lithium? Same # of valence e^-

16. Lewis structures are easier to draw than Bohr diagrams, but what information is lost by drawing a Lewis structure instead of a Bohr Diagram?

only valence e^-

Total # of e^-

17. Draw the Lewis structure for the following elements.

a) germanium — Ge

b) bromine — Br

c) xenon — Xe

d) potassium — K

e) arsenic — As

18. You should be able to tell how many valence electrons an atom has by which column of the periodic table the element is in. How many valence electrons are in each of the following atoms?

a) bromine — 7

b) tin — 4

c) krypton — 8

d) rubidium — 1

Summary of Info from POGIL #1

Tell me what you learned from this activity. Summarize in a list or a few short sentences.

Valence electrons determine how many dots to draw in Lewis Dot Structures. This # can be found using the column location on the periodic table.

POGIL #1 Extension

Bonding Sites: The number of **bonding sites** and atom has is the number of electrons it is willing to share in order to complete its octet.

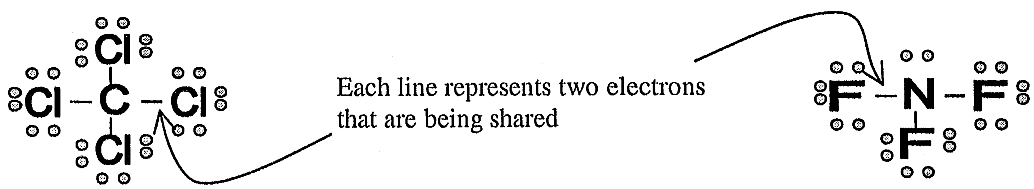
- Atoms with 4 valence electrons have __4__ bonding sites.
- Atoms that have 5 valence electrons have 3 bonding sites.
- Atoms with 6 valence electrons have __2__ bonding sites.
- Atoms with 7 valence electrons have __1__ bonding sites.

The exception to the rule is hydrogen. This element only has a 1s energy level, therefore it will be stable with __2__ valence electrons. Since hydrogen already has 1 electron, how many bonding sites will it have? __1__

POGIL #2 – Lewis Structures with Molecules

Information: Drawing Covalent Compounds

For covalent bonding, we often want to draw how the atoms share electrons in the molecule. For example, consider CCl_4 and NF_3 as drawn below:



Notice that the atoms share electrons so that they all have 8 electrons. If you count the electrons around carbon, you will get a total of eight (each line is two electrons). If you count the electrons around each chlorine atom, you will find that there are eight of them.

Critical Thinking Questions

1. How many valence electrons does a carbon atom have (before it bonds)? Hint: find this based on carbon's column on the periodic table. 4

2. How many valence electrons does a chlorine atom have (before it bonds)? 7

3. Since CCl_4 is made up of one carbon and four chlorine atoms, how many total valence electrons does CCl_4 have? Hint: add your answer to question 1 and four times your answer to question 2. C = 4 Cl = 7 (4) $4 + 4 \cdot 7 = 4 + 28 = 32$

4. Verify that there are 32 electrons pictured in the drawing of CCl_4. Yes

5. Find the sum of all the valence electrons for NF_3. (Add how many valence electrons one nitrogen atom has with the valence electrons for three fluorine atoms.)
   N = 5 F = 7 (3) $5 + 7 \cdot 3 = 5 + 21 = 26$

6. How many electrons are pictured in the drawing of NF_3 above?

26

7. In CCl_4 carbon is the "central atom". In NF_3 nitrogen is the "central atom". What is meant by "central atom"?

It is located in the center of the Lewis Dot Diagram

8. In SF_3 sulfur is the central atom. You can tell which atom is the central atom simply by looking at the formula. How does the formula give away which atom is the central atom?

There is only one sulfur so it is the central atom.

9. Identify the central atom in each of the following molecules:

A) CO_2 — Carbon (C)

B) PH_3 — P

C) SiO_2 — Si

10. For each of the compounds from question 9, add up how many valence electrons should be in the bonding picture. A is done for you.

A) CO_2 — **4+2 (6) =16**

B) PH_3 — $5+(1\cdot3)=8$

C) SiO_2 — $4+(2\cdot6)=16$

11. The number of electrons that should appear in the bonding picture for CO_3 is 22. The number of electrons that appear in the picture for CO_3^{2-} is 24. Offer an explanation for why CO_3^{2-} has 24 electrons instead of 22. (Where did the extra two electrons come from?)

CO_3^{2-} charge 2- means it gained 2 electrons

CO_3 is 22 + 2 more is $24e^-$

12. The number of electrons that should be included in the picture of NH_4 is 9. The number of electrons in the picture for NH_4^+ is 8. Offer an explanation for why NH_4^+ has 8 electrons instead of 9.

NH_4^{+} charge +1 means it lost 1 electron 9-1=8

13. Considering questions 11 and 12, we can formulate a rule: For each negative charge on a polyatomic ion, we must ADD (add or subtract) an electron and for each positive charge we must Subtract (add or subtract) an electron.

14. For each of the polyatomic ions or molecules below, determine the total number of valence electrons.

a) NO_3^- $5+3(6)+1=$ $5+18+1=24$

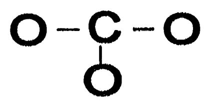
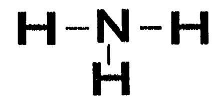
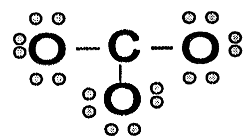
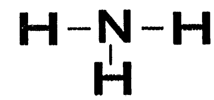
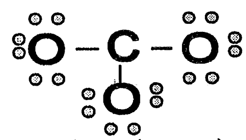
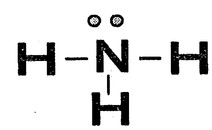
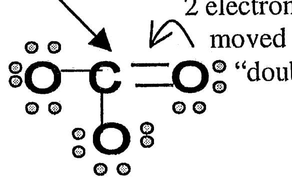
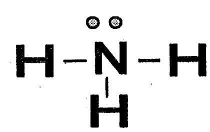
b) SCl_4 $6+4(7)=$ $6+28=34$

c) H_3O^+ $3(1)+6-1$ $3+6-1=8$

d) PO_4^{3-} $5+4(7)+3$ $5+28+3=36$

Information: Steps for Drawing Lewis Structures for Covalent Compounds

Study the two examples in the table of how to write structures for CO_3^{2-} and NH_3. Make sure you understand each of the five steps.

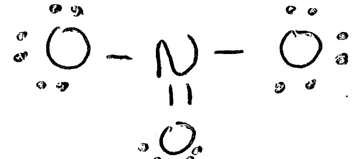
	CO_3^{2-}	NH_3
Step #1: Add up the number of valence electrons that should be included in the Lewis Structure.	4 + 3(6) + 2 = **24** (carbon has four and each oxygen has six; add two for the -2 charge)	5 + 3(1) = **8** (nitrogen has five; each hydrogen has one)
Step #2: Draw the "skeleton structure" with the central atoms and the other atoms, each connected with a single bond.		
Step #3: Add six more electron dots to each atom *except* the central atom. Also, never add dots to hydrogen.		 (no change)
Step #4: Any "leftover" electrons are placed on the central atom. Find the number of leftovers by taking the total from Step #1 and subtracting the number of electrons pictured in Step #3.	24 – 24 = 0 leftover electrons  (no change)	8 – 6 = 2 leftover electrons; placed around nitrogen 
Step #5: If the central atom has 8, then you are done. If not, then move two electrons from a different atom to make a multiple bond. Keep making multiple bonds until the central atom has 8 electrons.	a total of 4 electrons are shared here; 2 electrons were moved to form a "double bond" 	(no change) 

Critical Thinking Questions

15. Write the Lewis Structure for nitrate, NO_3^{-1}. Hint: when you are done it should look very similar to CO_3^{2-} in the table above.

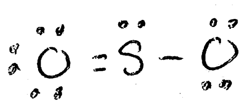
5 + 3(6) + 1

5 + 18 + 1 = 24

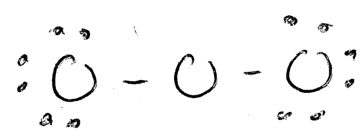
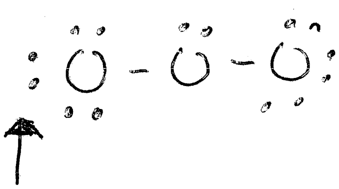
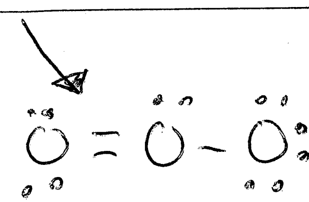


16. Draw the Lewis Structure for SO_2.

6 + 2(6) = 18



Summary of Info from POGIL #2

	Procedure	Example – O_3
Step #1	Add up the total valence electrons	$O_3 = 6 \cdot 3$ 18
Step #2	Draw you symbols and place a single bond between each	O - O - O
Step #3	Add 6 more electrons to everything <u>except</u> the <u>central atom</u>	
Step #4	Place left over electrons on the central atom.	 Move
Step #5	If central atom has 8 you are done! If not move 2 electrons to make a double bond (until central has 8)	

Molecular Shape NOTES

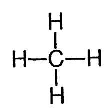
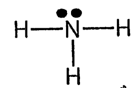
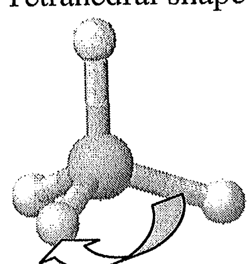
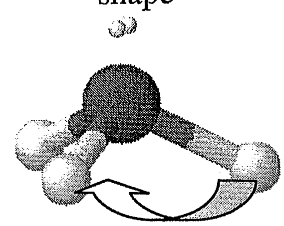
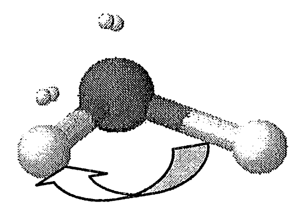
VSEPR Theory – A system for predicting the shape of molecules based on the idea that electrons orient themselves as far away as possible.

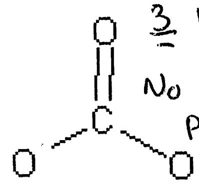
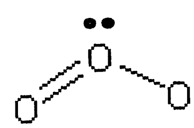
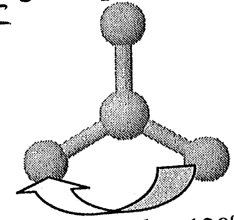
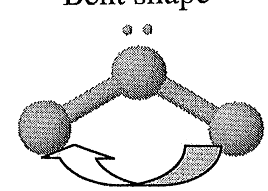
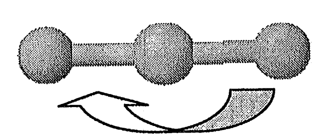
V- valence
S- shell
E- electrons
P- pair
R- repulsion

Why would electron pairs orient themselves as far apart as possible?

Same charges repel each other.

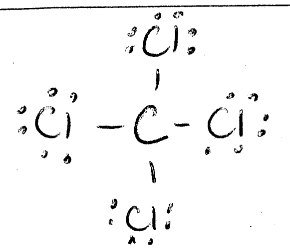
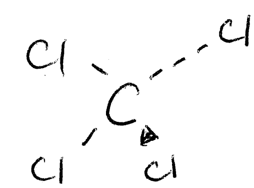
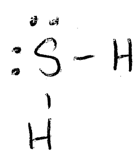
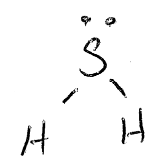
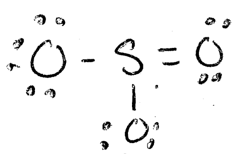
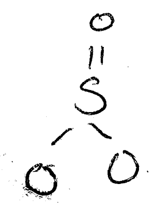
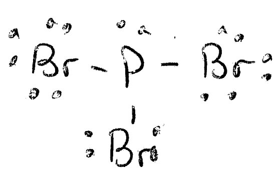
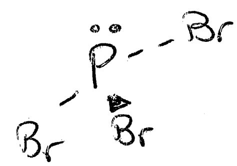
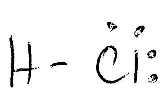
Information: Shapes of Molecules

Name	Methane, CH_4	Ammonia, NH_3	Water, H_2O
Lewis Structure	 4 bonds No lone pair	 3 bonds 1-lone pair	 2 bonds + lone pairs
3-D Shape (Around central atom)	Tetrahedral shape  Bond angle =109.5°	Trigonal pyramidal shape  Bond angle =106.5°	Bent shape  Bond angle =104.5°
Total # of electron regions	4	4	4
# of Bonding electron regions	4	3	2
# of lone pair electron regions	0	1	2

Name	Carbonate, CO_3^{2-}	Ozone, O_3	Carbon dioxide, CO_2
Lewis Structure	 3 bonds No lone pair	 2 bonds + lone pairs	O═C═O 2 bonds No lone pair
3-D Shape	Trigonal planar shape  Bond angle =120°	Bent shape  Bond angle =118.6°	Linear shape  Bond angle =180°
Total # of electron regions	3	3	2
# of bonding electron regions	3	2	2
# of lone pair electron regions	0	1	0

Steps for drawing molecular shapes

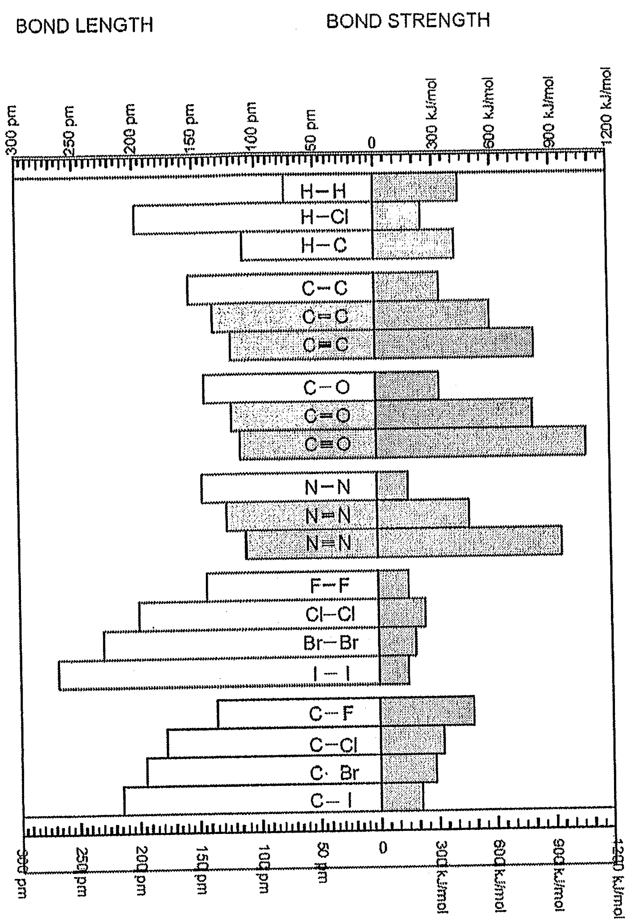
1. Draw the correct Lewis diagram for the molecule
2. Count the number of electron clouds around the central atom.
   a. Any type of bond (single, double, or triple) is one electron cloud.
3. Shape is based on number of electron clouds
   a. Two electron clouds = linear shape
   b. Three electron clouds = trigonal planar
   c. Four electron clouds = tetrahedral (also trigonal pyramidal)

	Lewis Structure	Number of SHARED pairs around central atom	Number of UNSHARED pairs around central atom	Molecular geometry	Drawing of molecule
CCl_4 $4+4(7)$ $4+28 = 32$		4	0	Tetrahedral	
H_2S $1(2)+6$ 8		2	2	Bent	
SO_3 $6+6(3)$ 24		3	0	Trigonal Planar	
PBr_3 $5+3(7)$ 26		3	1	Trigonal Pyramidal	
HCl $1+7$ 8		1	3	Linear	H -Cl

POGIL #3 – Relationships in Bonding

Part I:

BOND LENGTH AND STRENGTH



1. Which type of bond is the strongest: single, double, or triple bonds? Triple

2. Describe what you think may affect how strong a bond is between atoms.

- length of bond
- type of bond

3. Which type of bond is the longest: single, double, or triple bonds?

Single

4. Describe what you think may affect the length of bond between atoms.

Type of bond (attraction force/electronegativity)

5. State the relationship between bond strength and bond length. List any exceptions that you see.

Shorter the length the stronger the bond exceptions: H-H, H-Cl, H-C

Part II:

Each person needs to obtain one wooden splint. Trace the length of the wooden splint below.

6. Tracing of large splint:

Putting your hands at the outer edges of the splint, notice how much energy it takes to break the stick in half. Take one half of the stick and trace its length below. Putting your hands at the outer edge of the smaller splint, notice how much energy it takes to break the smaller splint in half.

7. Write a sentence comparing the energy it took to break the larger splint to the energy it takes to break the smaller splint.

You needed more energy to break the smaller flint.

8. Tracing of smaller splint:

The splints are a model for bond length and breaking them is a model for bond energy. Now, put two wooden splints together and notice how much energy it takes to break both sticks. Repeat with three wooden splints.

9. How does this relate to the conclusions you made on page 10?

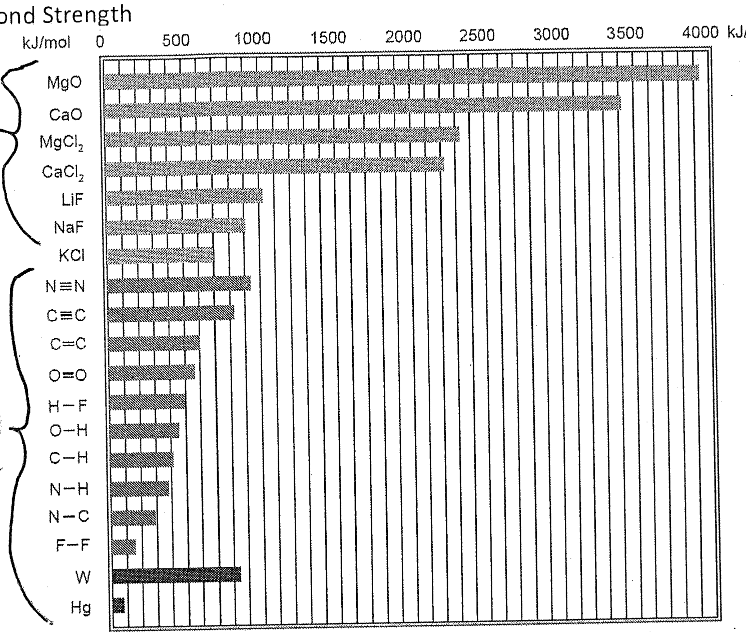
The more bonds, the more energy required to break them.

Discard your wooden stick pieces in the trash receptacle nearest you.

Part III:

STRONG AND WEAK BONDS

Bond Strength



Ionic

ovalent

On the graph to the left, draw a line separating ionic bonds from covalent bond.

10. According to the graph, in general, what type of bond is stronger, ionic or covalent bonds?

Ionic

11. Describe the relationship between the charge of the ion and strength of bond in ionic compounds.

The bigger the charge the stronger the bond.
(+1, +2, +3 etc.)

12. Why do you think this relationship exists?

Stronger, pulls it closer, harder to break

13. Which is stronger, single, double or triple bonds?

Triple

Summary of POGIL #3

14. Summarize what you have concluded regarding the ***strength of forces of attraction*** between molecules in compounds and *types of bonds*.

Strong attraction	Weak attraction
Triple bonds Ionic bonds (greater charge)	Single bonds Ionic bonds (lower charge)
Ionic is stronger than covalent	

a molecular shape formed when a central atom is bonded to three other atoms with no lone pairs of electrons on the central atom	a system for predicting the shape of molecules base on the idea that electrons orient themselves as far away as possible	a molecular shape in which two atoms are bonded to a central atom and no lone pairs exist on the central atom	an area of the electron cloud that will accept another electron and form a bond
a molecular shape in which one or two unshared pairs of electrons cause repulsion of two bonded atoms out of the same plane as the central atom	the distance between two bonded nuclei	a molecular shape formed when a central atom is bonded to four other atoms	a covalent bond in which three pairs of electrons are shared between two atoms
a covalent bond in which two pairs of electrons are shared between two atoms	two electrons in the outer shell of an atom that are not used in bonding	the energy required to break a bond	a covalent bond in which one pair of electrons are shared between two atoms

a depiction of an atom or molecule using the symbol to represent the nucleus and inner energy levels and dots or lines around the symbol to represent the number of bonds and valence electrons	a molecular shape formed when a central atom is bonded to three other atoms and a lone pair of electrons is found on the central atom	a pair of electrons involved in a bond between two atoms	