

Unit 1, Lesson 07: Answers to Homework on Covalent Bonding

1. Read pages 167 to 169.
2. Define bond energy (page 168). What does bond energy measure?
Bond energy is defined as the energy required to break the force of attraction between two atoms in a bond and separate them. Bond energy measures the amount of energy that is required to break a bond.
3. Which has higher bond energy: $C - C$, $C = C$, or $C \equiv C$?
 $C \equiv C$ has the highest bond energy. It takes the most energy to break a triple bond.
4. Do Questions 6, 7 and 8 on pages 169-170 of McGraw-Hill

Q6, p. 169: Rank the following compounds in sequence from lowest to highest melting point and give reasons for your decisions: $AsBr_3$, KBr , $CaBr_2$.

In general, the lower the ΔEN between the bonding atoms, the less polar the bond is, the weaker the inter-molecular attraction between molecules and the lower the melting point. So, in order from lowest to highest melting point: $AsBr_3$ ($\Delta EN = 0.78$) < $CaBr_2$ ($\Delta EN = 1.96$) < KBr ($\Delta EN = 2.14$)

Q7. p. 170: From their position in the periodic table, predict which bond in the following groups is the most polar. Verify your predictions by calculating ΔEN .

- a) $C - H$, $Si - H$, $Ge - H$: C , Si and Ge are all in the same group. Electronegativity decreases down a group, so $C - H$ will be the most polar bond because it has the larger ΔEN .
- b) $Sn - Br$, $Sn - I$, $Sn - F$: Br , I and F are all in the same group. Fluorine is the closest to the top of the periodic table, so it has the highest EN. $Sn - F$ will have the greatest ΔEN and will be the most polar bond.
- c) $C - O$, $C - H$, $C - N$: from its position on the periodic table, oxygen has the highest EN. The $C - O$ bond will be the most polar.

Q8, p. 170: Classify the bonding in each of the following as pure covalent, polar covalent or ionic. Afterwards, rank the polar covalent compounds in order of increasing polarity (use our criteria for pure and polar covalent bonds ($\Delta EN < 0.50$ and ΔEN between $0.50 - 1.70$, respectively))

- a) S_8 $\Delta EN = 0$ pure covalent
- b) $RbCl$ $\Delta EN = (3.16 - 0.82) = 2.34$ ionic
- c) PF_3 $\Delta EN = (3.98 - 2.19) = 1.79$ ionic
- d) SCl_2 $\Delta EN = (3.16 - 2.58) = 0.58$ polar covalent
- e) F_2 $\Delta EN = 0$ pure covalent
- f) SF_2 $\Delta EN = (3.98 - 2.58) = 1.40$ polar covalent

According to our criteria, both SCl_2 and SF_2 are polar. SCl_2 is less polar than SF_2 . I'm not sure why the text is classifying PF_3 as polar covalent. It may be because both P and F are non-metals, so the compound will have many covalent properties. Remember, our "cut-off" points for the ΔEN values are somewhat arbitrary.

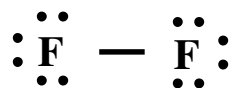
5. Complete the chart below to classify each bond as pure covalent, polar covalent or ionic. Use our criteria for pure and polar covalent bonds ($\Delta EN < 0.50$ and ΔEN between $0.50 - 1.70$, respectively):

Bonded Atoms	ΔEN	% Ionic Character	Type of Bond
a) H – Cl	$(3.16 - 2.20) = 0.96$	about 21%	polar covalent
b) Cl – F	$(3.96 - 3.16) = 0.80$	15%	polar covalent
c) Cl – Cl	$(3.16 - 3.16) = 0.00$	0%	pure covalent
d) Na – Cl	$(3.16 - 0.93) = 2.23$	about 71%	ionic
e) C – S	$(2.58 - 2.55) = 0.03$	essentially 0%	pure covalent
f) Cs – I	$(2.66 - 0.79) = 1.87$	between 58 – 59%	ionic
g) Fr – F	$(3.96 - 0.70) = 3.26$	about 92%	ionic
h) N – I	$(3.04 - 2.66) = 0.38$	almost 4%	pure covalent
i) F – O	$(3.96 - 3.44) = 0.52$	between 6 and 7%	polar covalent
j) C – O	$(3.44 - 2.55) = 0.89$	between 18 and 19%	polar covalent

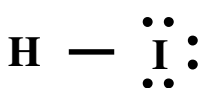
6. Use the Octet Rule to draw the structural diagrams (line diagrams) for the following covalent molecules. Recall the general rules from Grade 11:

- oxygen does not bond to oxygen except in O_2 , O_3 and H_2O_2
- if there are two or more atoms of carbon in a molecule, they are usually bonded to each other

a) F_2



b) HI



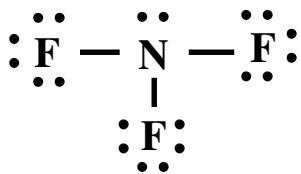
c) N_2



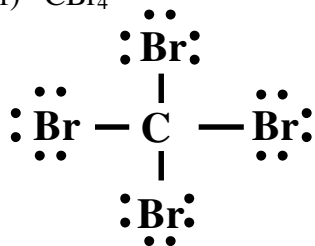
d) SiS_2



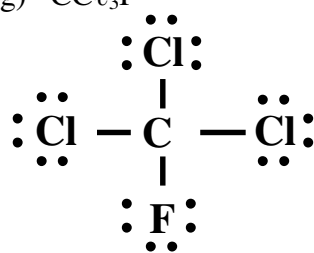
e) NF_3



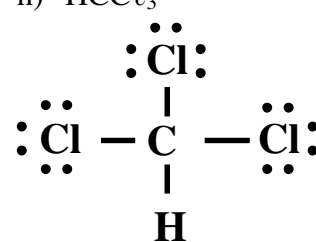
f) CBr_4



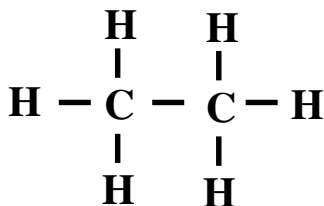
g) CCl_3F



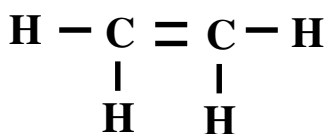
h) $HCCl_3$



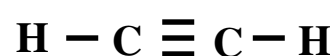
i) C_2H_6



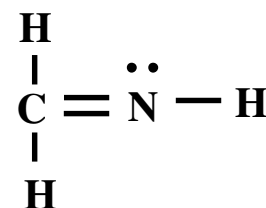
j) C_2H_4



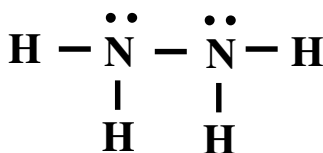
k) C_2H_2



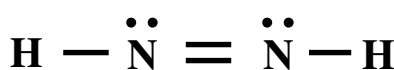
l) CH_2NH



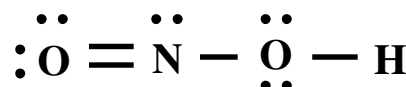
m) N_2H_4



n) N_2H_2

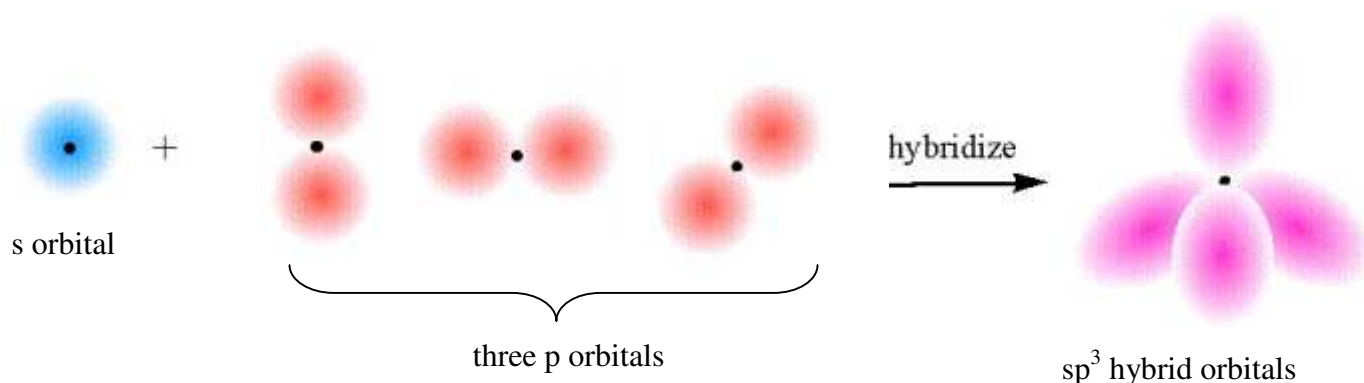


o) HNO_2

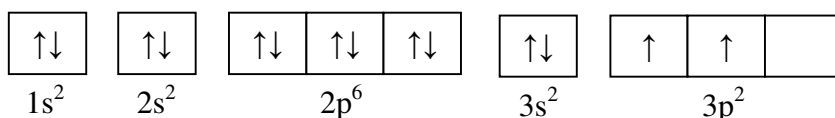


7. Explain what is meant by hybridized orbitals. Draw the predicted and actual orbital diagrams for silicon. Explain why the hybrid orbitals are named sp^3 .

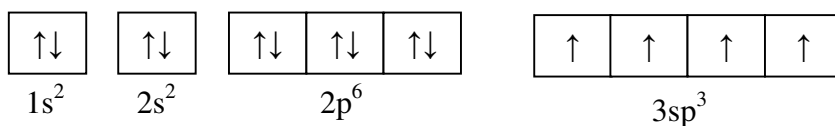
Hybrid orbitals are created when electrons from two different types of orbitals (eg. s and p) spread out equally around an atom. These electrons no longer travel in sphere-shaped or perpendicular orbitals. Rather, they have a new shape that is a blending (hybrid) of the s and p orbitals.



eg. silicon's predicted orbital diagram is:



silicon's actual orbital diagram is:



The hybrid orbitals are named sp^3 because they are formed when one "s" and three "p" orbitals are combined to form four equivalent orbitals.

8. In chart form, compare the ΔEN and typical physical properties of ionic and covalent compounds. How are these properties related to ΔEN ?

Property	Covalent Compounds	Ionic Compounds
ΔEN	0.00 to 1.70	> 1.70
State at SATP	gas, liquid or soft solid	solid
Melting and boiling points	very low to low	high to very high
Presence of odour	often have odours	usually do not have odours
Electrical conductivity	do not conduct in pure form	do not conduct in pure form
Electrical conductivity in solution	do not conduct in solution	conduct in solution (are electrolytes)

As ΔEN increases, the polarity of the bonds increases which means that the bond becomes more and more charged. As the charge increases, there is stronger inter-molecular attraction which holds the molecules together more tightly.

Because the molecules are more strongly attracted to one another:

- they are more likely to be liquid or solid at SATP
- they will have higher melting and boiling points
- they are less likely to have odours because there is less chance for particles to separate off and travel through the air into your nose

Pure ionic and covalent compounds do not conduct electricity because the valence electrons in these compounds are not “delocalized”. The non-metal atoms in both ionic and covalent compounds have high EN. This holds the valence electrons very close to the non-metal atom. They are not free to move from atom to atom and conduct an electric current.

9. Complete the following chart to compare the nature of the atoms that participate in the different types of bonding, and how this affects the properties of the substance they form:

	Metal Atoms IE is low EN is low	Non-metal Atoms IE is high EN is high
Metal Atoms IE is low EN is low	<p>Metallic Bonding (metal + metal)</p> <p>Neither bonding atom has a strong hold on the valence electrons so they are loosely held and free to move from atom to atom. The loose “sea of electrons” is responsible for the properties of metals.</p>	<p>Ionic Bonding (metal + non-metal)</p> <p>The non-metal strongly attracts the electron and pulls it from the metal. Fully charged ions are formed, causing very strong intermolecular attraction, which is responsible for the properties of ionic compounds.</p>
Non-metal Atoms IE is high EN is high	<p>Ionic Bonding (metal + non-metal)</p> <p>The non-metal strongly attracts the electron and pulls it from the metal. Fully charged ions are formed, causing very strong intermolecular attraction, which is responsible for the properties of ionic compounds.</p>	<p>Covalent Bonding (non-metal + non-metal)</p> <p>The non-metal atoms both have strong attraction for the bonding electrons, so they are held tightly between the bonded atoms. The bonding electrons will be shifted closer to whichever bonding atom has the higher EN.</p>