

Review for Unit Test #2: Chemical Bonding: Answers

Multiple Choice Questions:

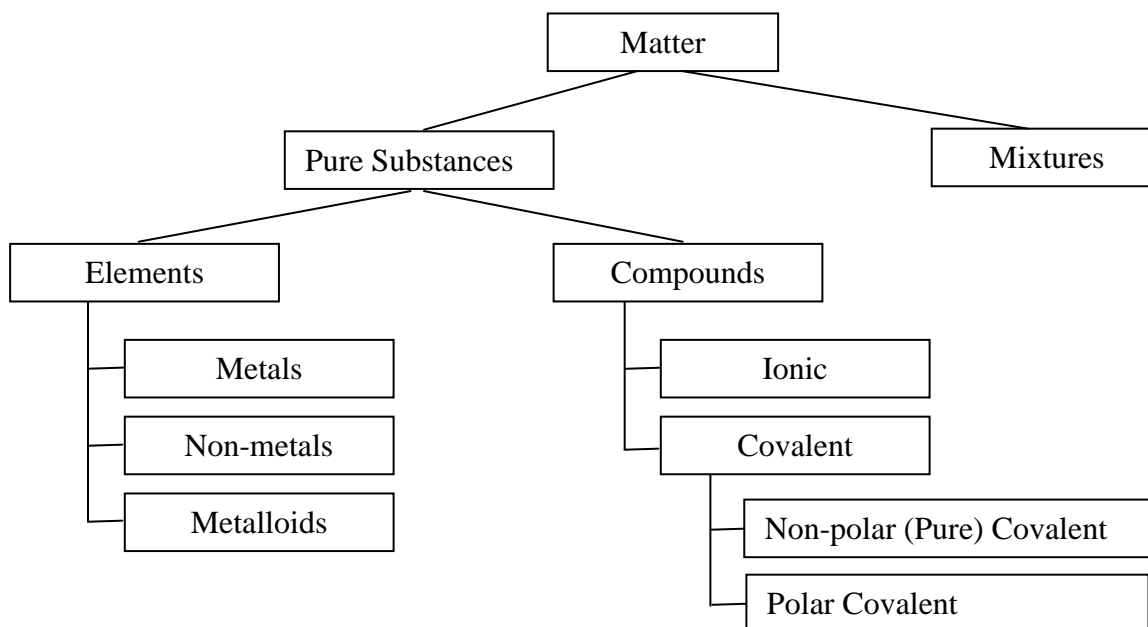
1. a	8. c	15. c	22. c	29. b	36. a
2. a	9. a	16. b	23. d	30. a	37. d
3. d	10. c	17. a	24. b	31. c	38. b
4. c	11. b	18. a	25. a	32. d	39. b
5. c	12. c	19. b	26. a	33. b	40. d
6. b	13. b	20. a	27. c	34. b	41. d
7. c	14. d	21. a	28. c	35. a	42. c

Short and Long Answer Questions:

1. Be able to define the following terms. Include one additional piece of information such as an example or application.

element	formula unit	molecular formula
compound	crystal lattice	intra-molecular forces
octet rule	electronegativity	inter-molecular forces
bonding capacity	non-polar (pure) covalent bond	hydrogen bonding
chemical bond	polar covalent bond	electrolyte
ionic bond	molecule	

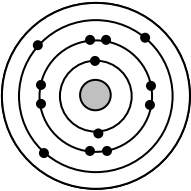
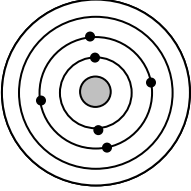
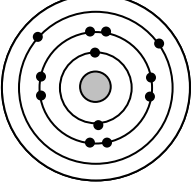
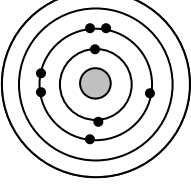
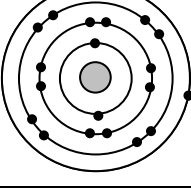
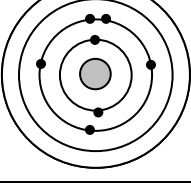
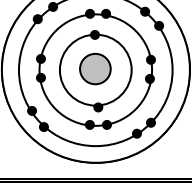
2. Complete the organization of matter chart. Include the terms: non-polar (pure) covalent compound, metal, compound, pure substance, polar covalent compound, ionic compound, metalloid, element, mixture, non-metal, covalent compound



3. Which of the following are pure substances: metals, non-metals, ionic compounds, covalent compounds, solutions? For each, explain why or why not.

Refer back to the organization of matter chart (above). Metals, non-metals, ionic compounds and covalent compounds are all pure substances because they contain only one type of particle. Because the atoms in compounds are chemically bonded together, they are considered to be particles. On the other hand, a solution is a type of mixture (for example, sugar in water), so it contains more than one type of particle, so solutions are mixtures.

4. Represent the following atoms three ways: using Rutherford-Bohr diagrams, electron configurations and Lewis Dot (electron dot) diagrams: Al, C, Mg, O, K, N and Ar

Element	Atomic #	Electron Configuration	Rutherford-Bohr Diagram	# of Valence Electrons	Electron Dot Diagram
Al	13	$1s^2 2s^2 2p^6 3s^2 3p^1$		3	$\cdot \overset{\cdot}{\underset{\cdot}{\text{Al}}}$
C	6	$1s^2 2s^2 2p^2$		4	$\cdot \overset{\cdot}{\underset{\cdot}{\text{C}}} \cdot$
Mg	12	$1s^2 2s^2 2p^6 3s^2$		2	$\cdot \text{Mg} \cdot$
O	8	$1s^2 2s^2 2p^4$		6	$\cdot \cdot \overset{\cdot \cdot}{\underset{\cdot}{\text{O}}} \cdot$
K	19	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$		1	$\text{K} \cdot$
N	7	$1s^2 2s^2 2p^3$		5	$\cdot \overset{\cdot \cdot}{\underset{\cdot}{\text{N}}} \cdot$
Ar	18	$1s^2 2s^2 2p^6 3s^2 3p^6$		8	$\cdot \cdot \overset{\cdot \cdot}{\underset{\cdot \cdot}{\text{Ar}}} \cdot \cdot$

5. Why are Lewis dot diagrams acceptable to use to show bonding?

Lewis dot diagrams show atoms with only their valence electrons. Because it is only the valence electrons that participate in bonding, these are the only electrons that need to be considered, and Lewis dot diagrams are acceptable to show bonding.

6. Explain why the Noble Gas elements do not form chemical bonds.

The Noble Gases all have a full outer electron shell: an s^2p^6 , or stable octet, electron arrangement. Because atoms are most chemically stable with a full outer shell, the Noble Gases will tend not to lose or gain electrons. This means that they will not form chemical bonds.

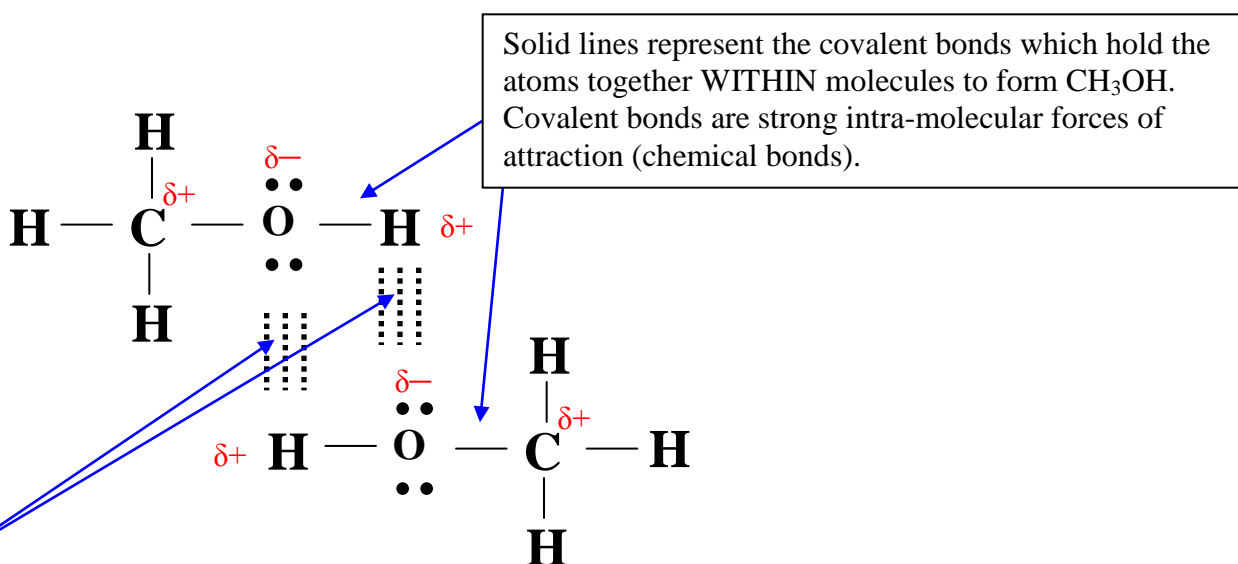
7. Complete the following table to compare the “typical” characteristics of ionic compounds and NON-POLAR (pure) covalent compounds.

Property	Ionic Compound	Non-polar (pure) Covalent Compound
Melting point	high or very high	very low or low
Hardness of solid	hard, crystalline solids	soft, waxy solids
Presence of odour	odourless	often have odours
Solubility in water	soluble	insoluble
Conduct electricity as a solid	do not conduct	do not conduct
Conduct electricity in solution	conduct in solution	do not conduct in solution

8. Clearly distinguish between inter-molecular and intra-molecular forces of attraction. See diagram below.

Intra-molecular forces of attraction are the attractions within a molecule that hold the atoms together. That is, intra-molecular attractions refer to chemical bonds (ionic and covalent).

Inter-molecular forces of attraction are the attractions between molecules. The more polar a molecule is, the more highly charged it will be and the stronger the attraction will be between the charged regions of adjacent molecules. Hydrogen bonding is an example of inter-molecular attraction.



The dashed lines represent the attraction between the charged regions of two molecules of CH_3OH . These are inter-molecular forces because they act BETWEEN molecules. For these two molecules, the inter-molecular attractions are also called “hydrogen bonds” because they are relatively strong. Technically, they are not bonds because they act between molecules, not within them and they are not as strong as actual covalent bonds. Hydrogen “bonding” occurs between molecules that have $\text{N}-\text{H}$ bonds, $\text{O}-\text{H}$ bonds or $\text{F}-\text{H}$ bonds because these bonds are extremely polar and have quite strong δ^+ and δ^- regions.

9. Use the strength of inter-molecular attraction to explain why:
- a) Ionic compounds have very high melting points

Ionic compounds are made up of fully charged positive and negative ions. Because these ions are all attracted to one another, they form a three-dimensional structure called a crystal lattice. In order to melt a crystal lattice, you must add enough energy to overcome the very strong attraction between the ions in the crystal lattice. Because ionic compounds have so many, and such strong, attractions between ions, it takes a great deal of heat to separate them so the substance will melt.

- b) Non-polar (pure) covalent compounds have low melting points and many are gases at room temperature.

The electrons in non-polar (pure) covalent bonds are shared equally between the bonded atoms. The electrons are not shifted significantly toward either atom, so there are no charged regions in the molecules. Because there are no charged regions, there is very little attraction between adjacent molecules, so it takes very little added energy (heat) to break the inter-molecular attraction. Therefore, non-polar compounds melt at low temperatures and many are gases at room temperature.

- c) Covalent compounds often have odours and ionic compounds do not.

Because of their low inter-molecular attractions, the molecules in covalent compounds can easily separate from one another and vapourize in the air. These vapourized molecules can land in your nose and be detected by “smell” receptors and this means that covalent compounds have odours.

Ionic compounds have very strong attraction between the ions in the crystal lattice (inter-molecular attraction), so it is unlikely that the ionic particles will separate from the crystal lattice and vapourize into the air. Because there are very few ionic particles in the air, ionic compounds can not land in your nose, so they have no odour.

10. Explain why ionic compounds are strong electrolytes in solution, and covalent compounds are not.

An **electrolyte** is defined as a substance that dissolves to produce ions in solution, which allows the solution to conduct an electric charge.

Ionic compounds are made of ions. When ionic compounds are dissolved in water, the partially charged regions of water molecules can “wobble in between” the ions and separate them. A solution of an ionic compound is really a solution of free ions surrounded by water molecules. If an electric current is passed through an ionic solution, the ions can conduct the electrons through the liquid, so ionic compounds are electrolytes.

Covalent compounds do not contain ions. Because they do not contain charged particles, they can not conduct an electric current through a solution, so they are not electrolytes.

11. Predict four physical properties of BaCl_2 . Because BaCl_2 is an ionic substance, it will:

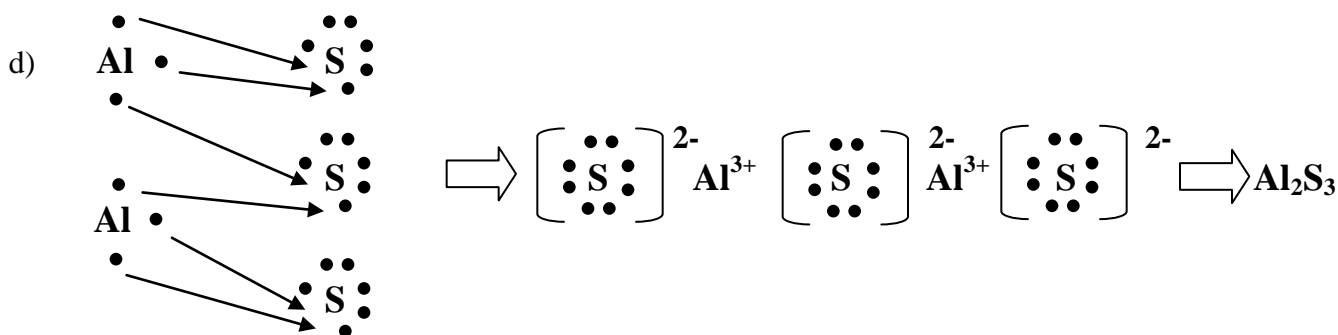
- form hard crystals at SATP
- have a high melting point
- be odourless
- not conduct electricity in pure, solid form
- dissolve well in water
- conduct electricity in solution (be an electrolyte)

12. Classify the following compounds as either **ionic** or **covalent** compounds:

- a) A compound has a strong odour, low melting point and does not dissolve in water: **covalent**
- b) A compound dissolves in water, has a high melting point and no odour: **ionic**
- c) $\text{C}_6\text{H}_{12}\text{O}_6$: **covalent**
- d) A compound is a liquid at room temperature and does not conduct an electric current: **covalent**
- e) MgCO_3 : **ionic**

- a) sodium and oxygen
b) barium and carbon
c) potassium and chlorine
d) aluminum and sulfur

Step 3: write the chemical formula for the final compound.



- Oxygen is double bonded and nitrogen is triple bonded.

15. Which of the following are impossible molecules? Why? Ar_2 , Ba_2 , NeO , He_2 , LiNa , HeH , Mg_3Al_2

All of these molecules are impossible. The Noble Gases will not form bonds because they already have a stable octet. Metals generally do not bond to metals to form individual molecules such as Mg_3Al_3 .

16. What **two** factors determine the electronegativity of an atom? (refer back to Atomic Theory Unit). Which element on the Periodic Table has the highest electronegativity, and why? Which element has the lowest, and why?

Both net nuclear attraction (NNA or Z_{eff}) and shielding effect determine the electronegativity of an atom.

As net nuclear attraction (NNA or Z_{eff}) increases going from left to right across a period, there is a stronger attraction between the atom's nucleus and the electrons in its valence shell. There is also a stronger attraction for another atom's electrons when one atom bonds to another atom to complete a stable octet electron arrangement. This stronger attraction for the electrons in a bond is called electronegativity.

As shielding effect increases down a group, electronegativity decreases. This is because the valence electrons are further from the atom's nucleus, so there is a weaker attraction for valence electrons and less of a pull on the electrons in a bond.

Electronegativity is highest at the top of the periodic table, to the far right (but not including the Noble Gases), so fluorine is the most electronegative. This is because fluorine has a very strong net nuclear attraction and a very small shielding effect, so it has a very strong attraction for the electrons in a bond.

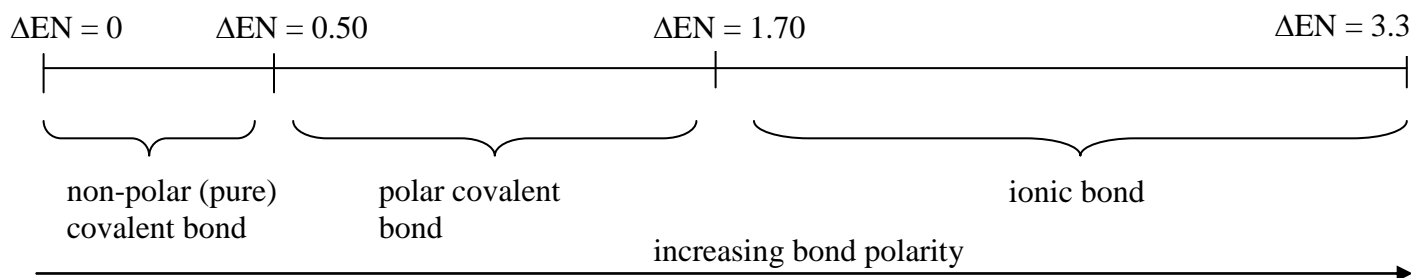
Electronegativity is lowest at the bottom of the periodic table, to the far left, so cesium (or francium) is the least electronegative. This is because cesium has a very small net nuclear attraction and a very large shielding effect, so it has a very weak attraction for the electrons in a bond.

17. Give three examples of bonds that have an electronegativity difference (ΔEN) of zero.

$\text{H}-\text{H}$, $\text{O}=\text{O}$, $\text{Br}-\text{Br}$, $\text{C}-\text{C}$, etc. Any bond between the same types of atoms will have $\Delta\text{EN} = 0$

18. Explain the concept of a "bonding continuum". Arrange the following bonds along a continuum from least to greatest polarity (increasing ΔEN): $\text{Na}-\text{Br}$, $\text{P}-\text{H}$, $\text{F}-\text{F}$, $\text{C}-\text{S}$, $\text{O}-\text{H}$, $\text{Cs}-\text{F}$, $\text{Ca}-\text{Cl}$, $\text{Mg}-\text{I}$. Which of these bonds would you expect to produce molecules with the strongest inter-molecular attraction? Which of these bonds will produce a compound with the highest melting point?

The bonding continuum:



The bonding continuum above shows that bonds may have a range of values for ΔEN from 0.0 to 3.30. That is, bond polarity can range anywhere from absolutely no polarity, to somewhat polar, to extremely polar (or ionic) bonds.

In order from least to most ionic, the bonds are:

$\text{F}-\text{F}$ ($\Delta\text{EN} = 0.00$)	$\text{P}-\text{H}$ ($\Delta\text{EN} = 0.01$)	$\text{C}-\text{S}$ ($\Delta\text{EN} = 0.03$)	$\text{O}-\text{H}$ ($\Delta\text{EN} = 1.24$)
$\text{Mg}-\text{I}$ ($\Delta\text{EN} = 1.35$)	$\text{Na}-\text{Br}$ ($\Delta\text{EN} = 2.03$)	$\text{Ca}-\text{Cl}$ ($\Delta\text{EN} = 2.16$)	$\text{Cs}-\text{F}$ ($\Delta\text{EN} = 3.28$)

As polarity increases, the bonding electrons are shifted more and more toward the bonding atom with higher electronegativity, creating more and more strongly charged regions on the molecule. If a molecule is asymmetrical, the charged regions on a molecule will be attracted to the oppositely charged regions on adjacent molecules and this is called inter-molecular attraction. The greater the ΔEN , the more strongly

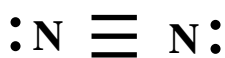
charged the molecules and the more strongly it will be attracted to other molecules. Therefore, the Cs – F bond, which has the greatest ΔEN will have the strongest inter-molecular attraction and it will also have the highest melting point.

19. Use the octet rule to draw the Lewis structures (aka stick diagrams or structural diagrams) for the following molecules. Remember to include all lone pairs (unshared electron pairs).

N_2 , NF_3 , N_2H_4 , N_2H_2 , CH_4 , PH_3 , C_2H_4 , CH_2O , CO_2 , CS_2 , $HSCN$, $HCOOH$, C_2H_5OH , $HSiN$, H_2CO_3 , CCl_3 , CH_3COCH_3 , CH_3COOH , $CHONH_2$, CH_3CONH_2 , OF_2 , HNO_2 , H_2O_2 , CH_3OH

- Write the AXE notation and name the molecular shape.
- Determine the symmetry of the molecule. Look at both the symmetry of the shape and the bonding atoms.
- Calculate the ΔEN of all bonds.
- Determine the overall polarity of the molecule. If the molecule is **very polar covalent** (asymmetrical and has polar bonds), then label partial negative (δ^-) and partial positive (δ^+) charges on the appropriate atoms. If the molecule is only slightly polar (asymmetrical but no polar bonds), the partial charges are very slight, so do not label them on the molecule.

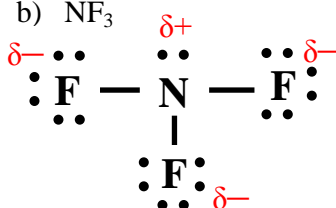
a) N_2



$$\Delta EN = 0.00$$

- no central atom (no AXE)
- linear
- symmetrical
- \therefore non-polar molecule
- \therefore no charges

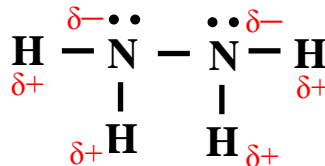
b) NF_3



$$\Delta EN = (3.98 - 3.04) = 0.94$$

- AX_3E_1
- trigonal pyramidal
- asymmetrical
- polar covalent bonds
- \therefore partial charges (label)

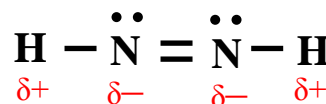
c) N_2H_4



$$\Delta EN = (3.98 - 3.04) = 0.94$$

- each N is AX_3E_1
- trigonal pyramidal
- asymmetrical
- polar covalent bonds
- \therefore partial charges (label)

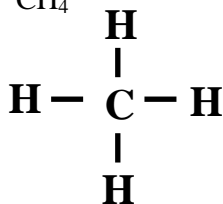
d) N_2H_2



$$\Delta EN = (3.98 - 3.04) = 0.94$$

- each N is AX_2E_1
- V-shaped or bent
- asymmetrical
- polar covalent bonds
- \therefore partial charges (label)

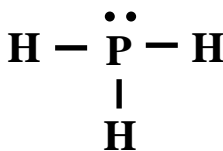
e) CH_4



$$\Delta EN = (2.55 - 2.20) = 0.35$$

- AX_4E_0
- tetrahedral
- symmetrical
- \therefore non-polar molecule
- \therefore no charges

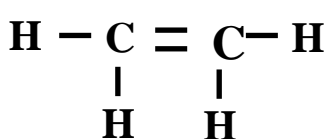
f) PH_3



$$\Delta EN = (2.20 - 2.19) = 0.01$$

- AX_3E_1
- trigonal pyramidal
- asymmetrical
- non-polar bonds
- \therefore tiny charges (don't label)

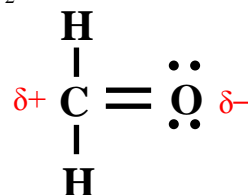
g) C_2H_4



$$\Delta EN = (2.55 - 2.20) = 0.35$$

- each C is AX_3E_0
- trigonal planar
- symmetrical overall
- \therefore non-polar molecule
- \therefore no charges

h) CH_2O



$$\Delta EN (\text{C-O}) = 3.44 - 2.55 = 0.89$$

- AX_3E_0
- trigonal planar
- asymmetrical (atoms)
- polar covalent bonds
- \therefore partial charges (label)

i) CO₂

$$\Delta\text{EN} = 3.44 - 2.55 \\ = 0.89$$

- AX₂E₀
- linear
- symmetrical
- ∴ non-polar molecule
- ∴ no charges

j) CS₂

$$\Delta\text{EN} = (2.58 - 2.55) \\ = 0.03$$

- AX₂E₀
- linear
- symmetrical
- ∴ non-polar molecule
- ∴ no charges

k) HSCN



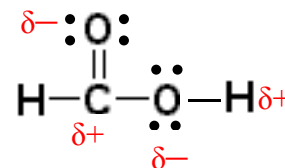
$$\Delta\text{EN} (\text{H}-\text{S}) = 0.38$$

$$\Delta\text{EN} (\text{S}-\text{C}) = 0.03$$

$$\Delta\text{EN} (\text{C}-\text{N}) = 0.49$$

- AX₂E₂ around sulfur (V-shaped or bent)
- AX₂E₀ around carbon (linear)
- asymmetrical (atoms)
- non-polar bonds
- ∴ tiny charges (don't label)

l) HCOOH

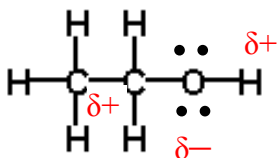


$$\Delta\text{EN} (\text{C}-\text{H}) = 0.35$$

$$\Delta\text{EN} (\text{C}-\text{O}) = 0.89$$

$$\Delta\text{EN} (\text{O}-\text{H}) = 1.24$$

- AX₃E₀ around carbon (trigonal planar)
- AX₂E₂ around oxygen (V-shaped or bent)
- asymmetrical
- polar covalent bonds
- ∴ partial charges (label)

m) C₂H₅OH

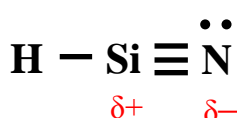
$$\Delta\text{EN} (\text{C}-\text{H}) = 0.35$$

$$\Delta\text{EN} (\text{C}-\text{O}) = 0.89$$

$$\Delta\text{EN} (\text{O}-\text{H}) = 1.24$$

- AX₄E₀ around carbons (tetrahedral)
- AX₂E₂ around oxygen (V-shaped or bent)
- asymmetrical
- polar covalent bonds
- ∴ partial charges (label)

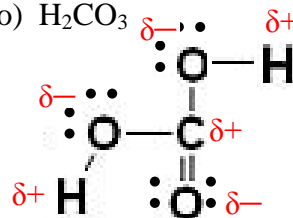
n) HSiN



$$\Delta\text{EN} (\text{H}-\text{Si}) = 0.30$$

$$\Delta\text{EN} (\text{Si}-\text{N}) = 1.14$$

- AX₂E₀
- linear
- asymmetrical (atoms)
- polar covalent bonds
- ∴ partial charges (label)

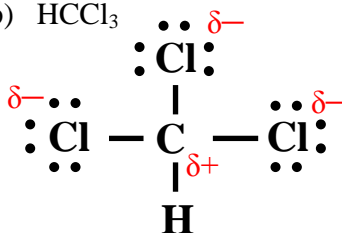
o) H₂CO₃

$$\Delta\text{EN} (\text{C}-\text{H}) = 0.35$$

$$\Delta\text{EN} (\text{C}-\text{O}) = 0.89$$

$$\Delta\text{EN} (\text{O}-\text{H}) = 1.24$$

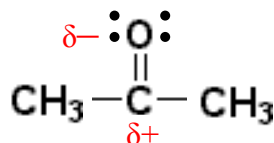
- AX₃E₀ around carbon (trigonal planar)
- AX₂E₂ around oxygen (V-shaped or bent)
- asymmetrical
- polar covalent bonds
- ∴ partial charges (label)

p) HCCl₃

$$\Delta\text{EN} (\text{C}-\text{H}) = 0.35$$

$$\Delta\text{EN} (\text{C}-\text{Cl}) = 0.61$$

- AX₄E₀
- tetrahedral
- asymmetrical (atoms)
- polar covalent bonds
- ∴ partial charges (label)

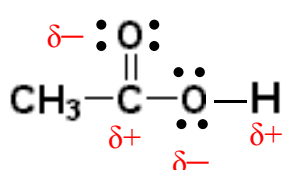
q) CH₃COCH₃

$$\Delta\text{EN} (\text{C}-\text{H}) = 0.35$$

$$\Delta\text{EN} (\text{C}-\text{O}) = 0.89$$

$$\Delta\text{EN} (\text{O}-\text{H}) = 1.24$$

- AX₃E₀ around central carbon (trigonal planar)
- asymmetrical atoms
- polar covalent bonds
- ∴ partial charges (label)

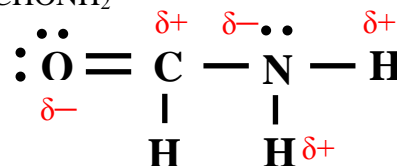
r) CH₃COOH

$$\Delta\text{EN} (\text{C}-\text{H}) = 0.35$$

$$\Delta\text{EN} (\text{C}-\text{O}) = 0.89$$

$$\Delta\text{EN} (\text{O}-\text{H}) = 1.24$$

- asymmetrical
- polar bonds
- ∴ partial charges (label)

s) CHONH₂

$$\Delta\text{EN} (\text{C}-\text{H}) = 0.35$$

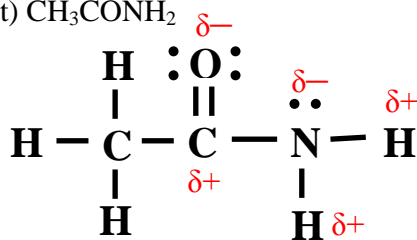
$$\Delta\text{EN} (\text{C}-\text{O}) = 0.89$$

$$\Delta\text{EN} (\text{N}-\text{H}) = 0.84$$

$$\Delta\text{EN} (\text{N}-\text{C}) = 0.49$$

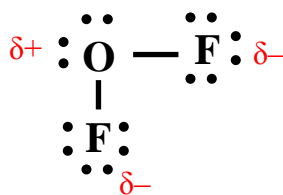
- asymmetrical
- polar bonds
- ∴ partial charges (label)

t) CH_3CONH_2



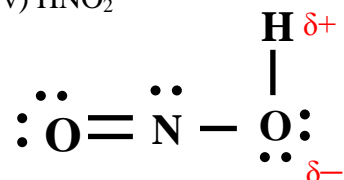
- $\Delta\text{EN} (\text{C} - \text{H}) = 0.35$
 $\Delta\text{EN} (\text{C} - \text{O}) = 0.89$
 $\Delta\text{EN} (\text{N} - \text{C}) = 0.49$
 $\Delta\text{EN} (\text{N} - \text{H}) = 0.84$
 · asymmetrical
 · polar bonds
 \therefore label partial charges

u) OF_2



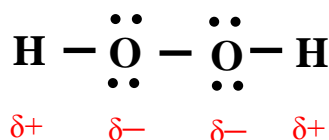
- $\Delta\text{EN} (\text{F} - \text{O}) = 0.54$
 · asymmetrical
 · polar bonds
 \therefore label partial charges
 · AX_2E_2
 · V-shaped (bent)

v) HNO_2



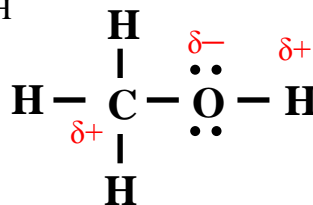
- $\Delta\text{EN} (\text{N} - \text{O}) = 0.40$
 $\Delta\text{EN} (\text{O} - \text{H}) = 1.24$
 · asymmetrical
 · polar bonds
 \therefore label partial charges
 · AX_2E_1 around N (bent)
 · AX_2E_2 around O (bent)

w) H_2O_2



- $\Delta\text{EN} (\text{O} - \text{H}) = 1.24$
 · AX_2E_2 around each O
 · V-shaped or bent
 · asymmetrical molecule
 · polar bonds
 \therefore label partial charges

x) CH_3OH



- $\Delta\text{EN} (\text{C} - \text{H}) = 0.35$
 $\Delta\text{EN} (\text{C} - \text{O}) = 0.89$
 $\Delta\text{EN} (\text{O} - \text{H}) = 1.24$
- AX_4E_0 around C (tetrahedral)
 · AX_2E_2 around O (bent)
 · asymmetrical molecule
 · polar bonds
 \therefore label partial charges

20. Explain why a molecule can contain polar bonds, but be a non-polar molecule.

If a molecule is symmetrical with regard to both its shape and bonded atoms, then the ΔEN for any polar bonds will act equally in opposite directions. This means that the electrons will not be shifted significantly in any direction (the partial charges will 'cancel out') so the molecule will actually be non-polar and have no partial charges.

21. Explain the relationship between the polarity of a molecule, the strength of inter-molecular attraction and melting point. Explain the relationship between the polarity of a molecule and the strength of inter-molecular attraction.

As the polarity of a molecule increases, the electrons are shifted away from the centre of a bond and toward the atom with the highest electronegativity. The further the electrons are shifted, the more charge they create on the atoms. The more strongly charged regions that there are within a molecule, the stronger the attraction will be BETWEEN molecules (inter-molecular attraction).