SOLUTIONS TO TOPIC H PROBLEMS

1) a) Cl_2S contains only nonmetals, so we expect it to be a **molecular** substance.

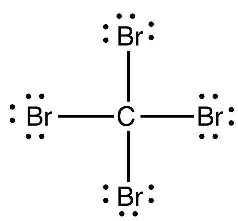
c) K_2S contains a metal and a nonmetal, so we expect it to be an **ionic** substance (containing K^+ and O^{2-} ions).

c) H_2SO_4 contains only nonmetals, so we expect it to be a **molecular** substance.

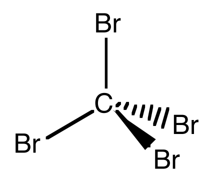
d) Fe is a metal, so it is a **metallic** substance.

e) He is a nonmetal, so it is a **molecular** substance. *(Yes, the noble gases contain individual atoms, not molecules, but they are still classified as molecular because their properties are similar to those of substance that do form molecules.)*

2) a) CBr_4 contains only nonmetals, so we expect it to be a molecular solid. When we draw the Lewis structure of CBr_4, we can deduce that the molecule is tetrahedral, which in turn means that this compound is nonpolar. Therefore, the only attractive forces between CBr_4 molecules are **London dispersion forces**.



Lewis structure of CBr_4



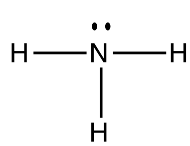
Actual shape of CBr_4

b) $CaBr_2$ is an ionic compound. The primary attractive force in any ionic compound is the **ion-ion attraction**. (Strictly speaking, the ion-ion attraction applies to Ca^{2+} and Br^- ions, rather than $CaBr_2$ units.)

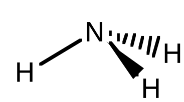
c) Ca is a metal. The attractive force between the atoms in any metal is **metallic bonding**, in which all of the atoms share their outer-shell electrons.

d) C is a network covalent solid in its two common forms (diamond and graphite). The attractive force between individual carbon atoms is **covalent bonds**.

3) a) NH_3 is a trigonal pyramidal molecule, so it is polar. It can also form hydrogen bonds, because the hydrogen atoms are directly bonded to nitrogen. Therefore, **all three types** of intermolecular forces are important for NH_3.

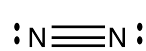


Lewis structure of NH_3



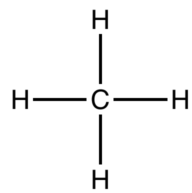
Actual shape of NH_3

b) N_2 is nonpolar and it cannot form hydrogen bonds, so only the **London dispersion force** is significant for N_2.

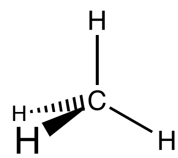


c) NF_3 is polar (like NH_3), but it cannot form hydrogen bonds. Therefore, the **London dispersion force** and the **dipole-dipole force** are significant for NF_3.

d) CH_4 is a tetrahedral molecule, so it is nonpolar. Its hydrogen atoms are not bonded to N, O, or F, so it cannot form hydrogen bonds. Therefore, only the **London dispersion force** is significant for CH_4.

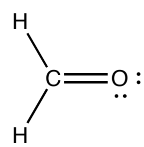


Lewis structure of CH_4



Actual shape of CH_4

e) CH_2O is a polar molecule (see the structure below), with the oxygen atom being negatively charged and the hydrogen atoms being positively charged. However, it cannot form hydrogen bonds, because its hydrogen atoms are not bonded to N, O or F. Therefore, the **London dispersion force** and the **dipole-dipole force** will be significant for CH_2O.



4) a) **BeO has a higher boiling point than CO**, because BeO is an ionic compound whereas CO is a molecular compound. Note that BeO contains a metal and a nonmetal, whereas CO contains two nonmetals.

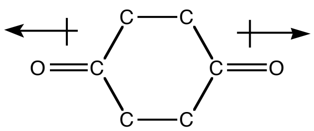
b) **MgO has a higher boiling point than NaF**. Both of these compounds are ionic, but the charges on the ions in MgO are larger than the charges in NaF (+2/-2 versus +1/-1). The higher the charges on the ions, the stronger the attraction between the ions, and this stronger attraction makes it more difficult to boil the compound.

5) a) HCl is a **gas**, because its boiling point is far below room temperature.

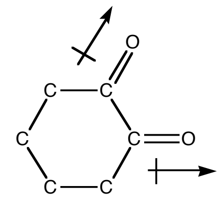
b) The melting and boiling points of HCl are much too low for an ionic compound. Furthermore, both H and Cl are nonmetals, so we would expect these elements to form a covalent bond. Based on this, we conclude that HCl is actually a **molecular compound**.

6) As we go from F to Cl to Br to I, the size and polarizability of the halogen atom increases. London dispersion forces are proportional to polarizability; the more polarizable the atom, the larger the instantaneous dipole. Therefore, the boiling point trend is a result of the increasing strength of the London dispersion forces.

7) Compound 1, because of its symmetry, is nonpolar; all of the C=O and C–H bond dipoles cancel out. Compound 2 is polar because the two C=O bond dipoles reinforce each other.



Compound 4: bond dipoles cancel out.

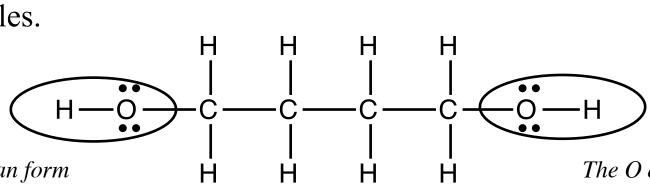


Compound 5: bond dipoles do not cancel out.

Both compounds have London dispersion forces, which should be nearly equal in strength, since the two compounds contain the same atoms. However, compound 2 has dipole-dipole attraction while compound 1 does not, so compound 2 has the stronger overall intermolecular attraction. This allows to assign the boiling points:

compound 1: boiling point = 70ºC
compound 2: boiling point = 116ºC

8) The difference here is the ability of each molecule to form hydrogen bonds. Compound 1 contains two –O–H groups, each of which can form hydrogen bonds to –O–H groups on neighboring molecules.



The O and the H can form hydrogen bonds to the H and the O of other molecules.

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Compound 2 contains only one –O–H group, while compound 3 contains none. Therefore, the order of strength of hydrogen bonding will be:

compound 1 > compound 2 > compound 3

The stronger the intermolecular forces, the higher the boiling point, so the correct answer is:

compound 1: boiling point = 235ºC
compound 2: boiling point = 135ºC
compound 3: boiling point = 82ºC

9) a) Ammonia is a gas at 25ºC and 1 atm. This point lies within the gas region of the diagram.

b) Ammonia is a solid at -80ºC and 10 atm. This point lies within the solid region of the diagram.

c) From the graph, we see that the triple point is at –77.75ºC and 0.060 atm.

d) From the graph, we see that the critical point is at 132.4ºC and 111.3 atm.

e) Ammonia will start as a solid, change to a liquid at some temperature slightly above –77.75ºC, and then change to a gas at some temperature between 25ºC and 132.4ºC.

f) Ammonia will start as a solid and change to a gas at some temperature between -100˚C and -77.75˚C. 0.01 atm is below the triple point pressure, so the ammonia will never be a liquid.

g) The ammonia will start as a gas, change to a liquid at some pressure slightly above 0.060 atm, and then change to a solid at some pressure above 0.060 atm. (Note that the temperature is just 0.05ºC above the triple point, so increasing the pressure to 150 atm is guaranteed to cross the liquid/solid boundary.)

h) 200 atm is higher than the critical pressure, so we will not observe any change of state! The ammonia will simply remain as an unusually compressible fluid throughout the heating, much more dense than a typical gas but somewhat less dense than "normal" liquid ammonia.

See below for an illustration of these answers.