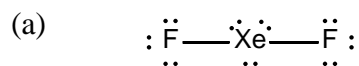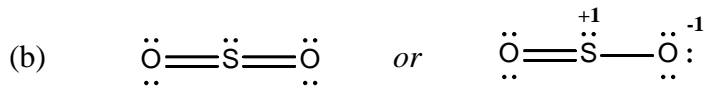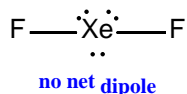# Answers to Practice Test Questions 9
Polarity, Intermolecular Forces, Kinetic Molecular Theory and Gases

1.

(a)

$: \ddot{F} — \ddot{Xe} — \ddot{F} :$

linear molecular geometry

$F — \ddot{Xe} — F$

no net dipole

(b)

$\ddot{O} = \ddot{S} = \ddot{O}$ or $\ddot{O} = \overset{+1}{S} — \ddot{O} :^{-1}$
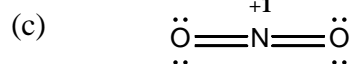
bent molecular geometry

net dipole

(c)

$\ddot{O} = \overset{+1}{N} = \ddot{O}$

linear molecular geometry

$O — N — O$

no net dipole

(d)
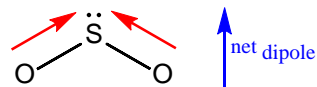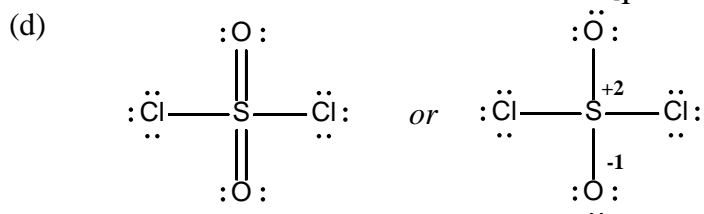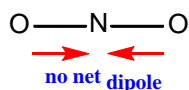
tetrahedral molecular geometry
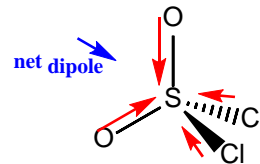
net dipole

2.   *All carbon atoms in all five molecules are trigonal planar (~120° bond angles).*

(a)

(b)

(c)
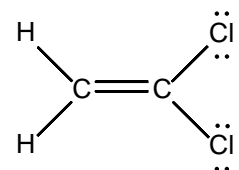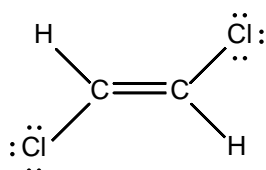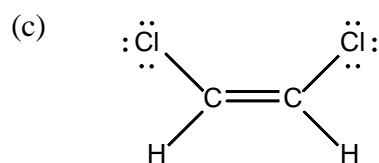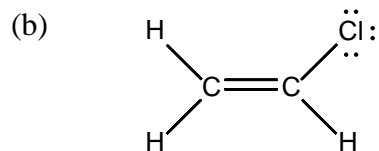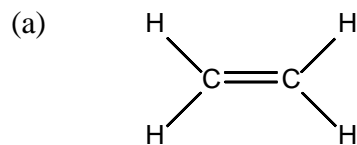
(d)   The polar molecules have net dipoles (shown below).
       The nonpolar molecules do not (since the bond dipoles all cancel).
       C_2H_4 is nonpolar.
       C_2H_3Cl is polar:
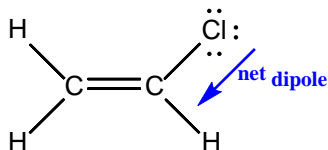
       The $C_2H_2Cl_2$ molecules on the left and right are polar:

       The $C_2H_2Cl_2$ molecule in the middle is nonpolar.

3.   Kevin didn't use VSEPR to determine the correct molecular geometry for CH_2F_2.
       CH_2F_2 is tetrahedral *not* square planar!
       As a result, the bond dipoles do *not* all cancel and there is a net dipole (as shown in the diagram below).
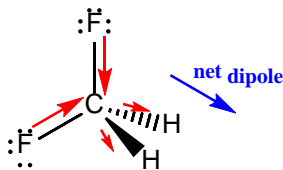
       *If Kevin had switched the places of one H and one F in his Lewis diagram, he would have correctly concluded that CH_2F_2 was polar *however* he would still have been skipping an essential step and he most certainly would *not* have got full marks! Fluking into the correct answer using methods that are not scientifically valid doesn't get credit.*

4.   *P-Cl bond dipoles should be drawn as an arrow pointing from Cl to P. P-F bond dipoles should be drawn as an longer arrow pointing from F to P. The length of the arrows should show that the P-F dipoles are larger than the P-Cl dipoles because the electronegativity difference is greater.*

(a)
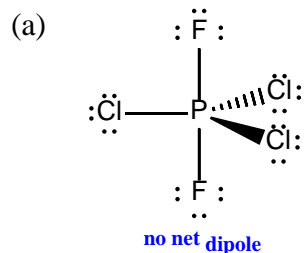
       Cl is larger than F therefore the Cl occupy the equatorial sites.
       The three P-Cl dipoles cancel. The two P-F dipoles cancel.
       There is therefore no net dipole and the molecule is nonpolar.

(b)
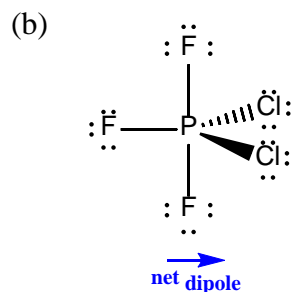
Cl is larger than F therefore Cl occupy two of the equatorial sites.

The two P-F$_{axial}$ dipoles cancel; however, the two P-Cl dipoles and the P-F$_{equatorial}$ dipole do not. (F is more electronegative than Cl, so the P-F dipole is larger.)

There is therefore a net dipole and the molecule is polar.

(c)
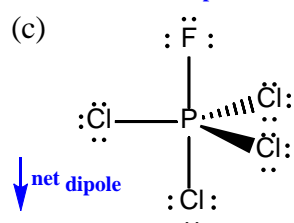
PFCl$_4$ has a net dipole so it is polar.

F is smaller than Cl therefore F occupies an axial site.

The three P-Cl$_{equatorial}$ dipoles cancel; however, the P-F dipole and the P-Cl$_{axial}$ dipole do not. (F is more electronegative than Cl, so the P-F dipole is larger.)

(d)
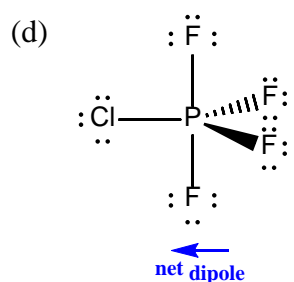
PF$_4$Cl has a net dipole so it is polar.

Cl is larger than F therefore Cl occupies an equatorial site.

The two P-F$_{axial}$ dipoles cancel; however, the P-Cl dipole and the two P-F$_{equatorial}$ dipoles do not. (F is more electronegative than Cl, so the P-F dipoles are larger.)

5. *In order to answer this question you have to draw the Lewis structures and determine the molecular geometries.*

(a) NaF consists of Na$^+$ cations and F$^-$ anions. Neither ion has a dipole moment.

Intermolecular forces:

Ion-Ion forces

Ion-induced dipole forces

Induced dipole-induced dipole forces

(b) H$_2$S has a bent molecular geometry and is polar.

Intermolecular forces:

Dipole-dipole forces

Dipole-induced dipole forces

Induced dipole-induced dipole forces
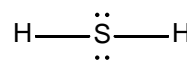
(c) SF$_6$ has an octahedral molecular geometry and is non-polar.

Intermolecular forces:

Induced dipole-induced dipole forces
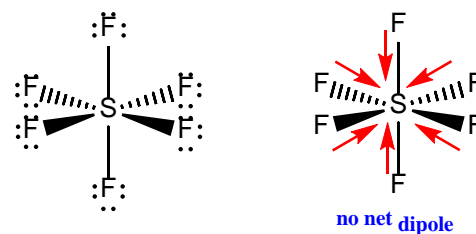
(d) $NaNO_3$ consists of Na^+ cations and NO_3^- anions. Neither ion has a dipole moment.
       Intermolecular forces:
              Ion-Ion forces
              Ion-induced dipole forces
              Induced dipole-induced dipole forces
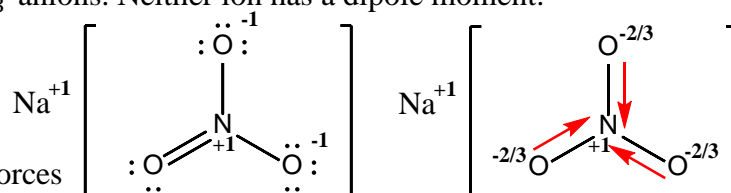
There are three resonance structures for NO_3^-. One is shown.
The VSEPR geometry picture accounts for all three resonance structures.

(e) SeF_4 has a seesaw molecular geometry and is polar.
       Intermolecular forces:
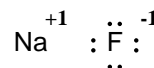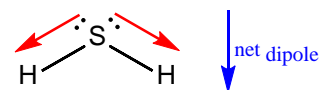   Dipole-dipole interactions
              Dipole-induced dipole interactions
              Induced dipole-induced dipole forces
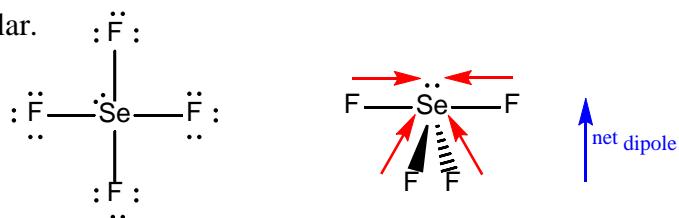
(f) OF_2 has a bent molecular geometry and is polar.
       Intermolecular forces:
              Dipole-dipole interactions
              Dipole-induced dipole interactions
              Induced dipole-induced dipole forces

6.   *Hydrogen bonding is an <u>intermolecular</u> attraction between atoms of different molecules.*
      It is <u>not</u> the same as covalent bonding!!!

(a) F^- anions can hydrogen bond with H in water

(b)   No atoms in H_2S can hydrogen bond with H or O in water.
      S is too large, and the H atoms are not positive enough.

(c)   F atoms in SF_6 can hydrogen bond with H in water.

(d) $O^{-2/3}$ atoms in NO_3^- can hydrogen bond with H in water.
      N in NO_3^- has no available lone pairs for hydrogen bonding *(see Lewis structure).*

7.   Both tetrahalides are tetrahedral, nonpolar compounds, so the only intermolecular forces
      present are induced dipole-induced dipole forces

      • CCl_4 is smaller than CI_4. Therefore, CI_4 has a much higher polarizability and has much
        stronger induced dipole-induced dipole forces. As a consequence, CI_4 has a higher
        boiling point than CCl_4.

      • The difference in strengths of the intermolecular forces results in different states of
        matter for these two compounds at room temperature. CCl_4 is a liquid while CI_4 is a
        solid.

8.

(a)
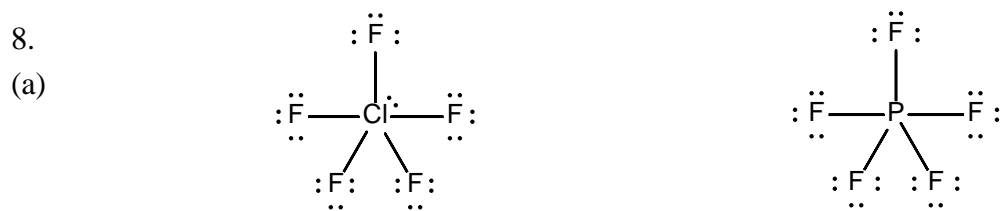
(b) ClF_5 dipole – dipole forces

PF_5 induced dipole – induced dipole forces (aka London dispersion forces)

(c) PF_5 is expected to have a lower boiling point than ClF_5.

For substances of comparable size, boiling point increases as the strength of intermolecular forces increases.

9.

net **dipole**

no net **dipole**
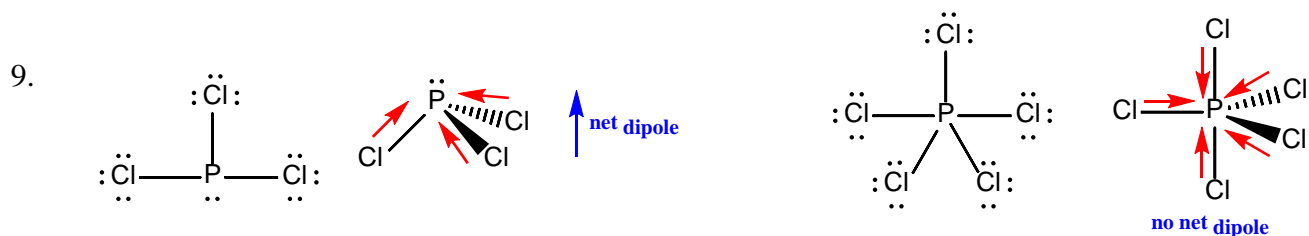
(a) PCl_3 is polar while PCl_5 is nonpolar. As such, the only intermolecular forces active in PCl_5 are induced dipole-induced dipole forces (London dispersion forces). In PCl_3, there are also dipole-dipole forces and dipole-induced dipole forces.

(b) PCl_5 is larger than PCl_3 so the induced dipole-induced dipole forces (London dispersion forces) in PCl_5 are stronger than in PCl_3.

(c) Since the boiling point of PCl_5 is higher than that of PCl_3, the increased strength of the induced dipole-induced dipole forces (London dispersion forces) must be greater than the strength of the "extra" intermolecular force types in PCl_3 (i.e. the dipole-dipole and dipole-induced dipole forces).

10.

(a) *lowest bp* CH_2O CH_3OH $NaOCH_3$ *highest bp*

Boiling point increases with the strength of the intermolecular forces (IMF) attracting the molecules to each other.

$NaOCH_3$ is an ionic compound, so its strongest IMF are ion-ion.

CH_2O and CH_3OH are polar, so their strongest IMF are dipole – dipole; however, CH_3OH can hydrogen bond while CH_2O cannot so its dipole – dipole forces should be stronger.

(b) *lowest bp* $Si(CH_3)_4$ $Ge(CH_3)_4$ $Sn(CH_3)_4$ *highest bp*

All three compounds are nonpolar so the only IMF to consider are induced dipole – induced dipole (London dispersion forces). The strength of these forces increases with the polarizability of the molecule which, in turn increases with size. Since $Sn(CH_3)_4$ is largest, its induced dipoles are largest and therefore its intermolecular forces are strongest (giving it the highest boiling point). Similarly, $Si(CH_3)_4$ is smallest, has the smallest induced dipoles, the weakest IMF and the lowest boiling point.

11.

(a)   Incorrect. A molecule can't have a temperature. Temperature is a macroscopic property.

(b)   Incorrect. Within a sample, the speeds of the gas molecules are distributed according to a Maxwell-Boltzmann distribution.

(c)   Incorrect. A molecule can't have a temperature. Temperature is a macroscopic property.

(d)   Correct.

12.

(a) F_2. It has a smaller molar mass.

(b) CH_4. It has a smaller molar mass.

(c) O_2. It has a smaller molar mass.

13.   **Step 1: Convert temperature into Kelvin**

$T = 25 + 273.15 = 298.15\ K = 298\ K$

This value only has 3 sig. fig. because 25 °C has no decimal places; however, we will still use 298.15 K in the calculation to minimize rounding error.

Step 2: Calculate average kinetic energy for one mole of gas

$\overline{E_k} = \frac{3}{2}RT$

$\overline{E_k} = \frac{3}{2}\left(8.314\ 462\ \frac{J}{mol \cdot K}\right)(298.15\ K)$

$\overline{E_k} = 3.72 \times 10^3\ \frac{J}{mol}$

Step 3: Calculate average kinetic energy for one molecule of gas

$\overline{E_k} = 3.72 \times 10^3\ \frac{J}{mol} \times \frac{1\ mol}{6.022141 \times 10^{23}\ molecules} = 6.17 \times 10^{-21}\ \frac{J}{molecule}$

Step 4: Check your work

Does your answer seem reasonable? Are sig. fig. correct?

The energy per molecule should be *much* smaller than the energy per mole since there are many molecules in a mole.

14.   **Step 1: Calculate molar mass in kg/mol**

$M_{UF_6} = 238.029\ \frac{g}{mol} + 6\left(18.9984\ \frac{g}{mol}\right) = 352.019\ \frac{g}{mol} \times \frac{1\ kg}{1000\ g} = 0.352019\ \frac{kg}{mol}$

Step 2: Calculate root-mean-square speed for each gas

$v_{rms} = \sqrt{\frac{3RT}{M}}$

$v_{rms}(Ar) = \sqrt{\frac{3\left(8.314462 \frac{J}{mol \cdot K}\right)(298K)}{\left(0.352019 \frac{kg}{mol}\right)}} \times \frac{1 \frac{kg \cdot m^2}{s^2}}{1J} = 145\ \frac{m}{s}$

Step 3: Check your work

Does your answer seem reasonable? Are sig. fig. correct?

15. **Step 1: Convert temperature into Kelvin**

$T = 0 + 273.15 = 273.15\ K = 273\ K$

This value only has 3 sig. fig. because 0 °C has no decimal places; however, we will still use 273.15 K in the calculation to minimize rounding error.

Step 2: Calculate molar mass in kg/mol

$M_{UF_6} = 2\left(15.9994\frac{g}{mol}\right) = 31.9988\frac{g}{mol} \times \frac{1\ kg}{1000\ g} = 0.0319988\ \frac{kg}{mol}$

Step 3: Calculate root-mean-square speed for each gas

$v_{rms} = \sqrt{\frac{3RT}{M}}$

$v_{rms}(Ar) = \sqrt{\frac{3\left(8.314462\frac{J}{mol\cdot K}\right)(273.15K)}{\left(0.0319988\frac{kg}{mol}\right)}} \times \frac{1\frac{kg\cdot m^2}{s^2}}{1J} = 461\frac{m}{s}$

Step 4: Check your work

Does your answer seem reasonable? Are sig. fig. correct?

16.

(a) **Step 1: Convert volume into m³**

$V = 500\ mL \times \frac{1\ L}{1000\ mL} \times \frac{1\ m^3}{1000\ L} = 5.00 \times 10^{-4}\ m^3$

Step 2: Calculate pressure if gas behaves ideally

$PV = nRT$

$P = \frac{nRT}{V} = \frac{(5.00\ mol)\left(8.314\ 462\frac{Pa\cdot m^3}{mol\cdot K}\right)(300\ K)}{5.00\times10^{-4}m^3} = 2.49 \times 10^7 Pa \times \frac{1\ bar}{10^5\ Pa} = 249\ bar$

Step 3: Check your work

Does your answer seem reasonable? Are sig. fig. correct?

This is a really high pressure, but it's a lot of gas in a small container. Maybe reasonable. See part (b)

(b) **Step 2: Calculate pressure if gas behaves nonideally**

$\left(P + a\frac{n^2}{V^2}\right)(V - b \cdot n) = nRT$

$P = \frac{nRT}{V-b\cdot n} - a\frac{n^2}{V^2}$

$P = \frac{(5.00\ mol)\left(8.314\ 462\frac{Pa\cdot m^3}{mol\cdot K}\right)(300\ K)}{(5.00\times10^{-4}m^3)-\left(3.20\times10^{-5}\frac{m^3}{mol}\right)(5.00\ mol)} - \left(0.1355\frac{Pa\cdot m^6}{mol^2}\right)\frac{(5.00\ mol)^2}{(5.00\times10^{-4}m^3)^2}$

$P = (3.67 \times 10^7 Pa) - (1.355 \times 10^7 Pa)$

$P = 2.31 \times 10^7 Pa \times \frac{1\ bar}{10^5\ Pa} = 231\ bar$

Step 3: Check your work

Does your answer seem reasonable? Are sig. fig. correct?

This is a really high pressure, but it's a lot of gas in a small container so it might be reasonable. This pressure is lower than that calculated by the ideal gas law – which is what we expect for most gases. (Exceptions being very small and nonpolar like He.)

(c)   The pressure calculated using the van der Waals equation should be closer to the actual pressure. The van der Waals equation accounts for intermolecular forces between gas particles as the density of the gas increases (which is why the pressure calculated in part (b) was lower) as well as the fact that, at high density, the volume occupied by the actual particles of gas is not negligible.

17.   HCl

The pressure calculated using the van der Waals equation should be closer to the actual

HCl is polar while H_2 is nonpolar *and* H_2 is not significantly larger than HCl. (In fact, it's smaller.) As such, the strongest intermolecular forces (dipole – dipole forces) between HCl molecules will be stronger than the strongest intermolecular forces (induced dipole – induced dipole forces) between H_2 molecules. Since a corrects for the intermolecular forces between gas molecules, HCl is expected to have a larger value for a.

18.

(a) Br_2

The van der Waals constant a accounts for the intermolecular forces between the gas particles.

All three molecules are nonpolar, so the only intermolecular forces to account for are induced dipole-induced dipole forces. The strength of these forces increases as the size of the molecule increases, so they should be strongest in Br_2. As such, Br_2 has the highest van der Waals constant a.

(b) Br_2

The van der Waals constant b accounts for the volume occupied by the gas particles.

Larger molecules occupy more volume than the same number of smaller molecules, so the largest molecule (Br_2) has the highest van der Waals constant b.

19.

(a)   NaCl(s) consists of Na^+ cations and Cl^- anions. Neither ion has a dipole moment.

Intermolecular forces:

Ion-Ion forces

Na^{+1} $:\overset{..}{\underset{..}{Cl}}:^{-1}$

Ion-induced dipole forces

Induced dipole-induced dipole forces

(b) H_2O has a bent molecular geometry and is polar. It can also hydrogen bond.

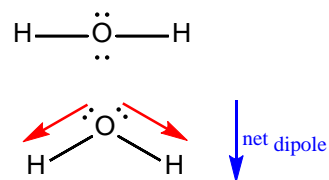   Intermolecular forces:

            Hydrogen bonding

            Dipole-dipole forces

            Dipole-induced dipole forces

            Induced dipole-induced dipole forces

(c)   When NaCl(s) dissolves in water, the intermolecular forces between the ions must be overcome. This requires energy (primarily to overcome the ion-ion forces since they are the strongest).

       Intermolecular forces between some of the water molecules must also be overcome (as the ions become solvated). This also requires energy.

       Energy is released, however, when new intermolecular forces become active between the water molecules and the Na^+ and Cl^- ions. The strongest new intermolecular forces are the ion-dipole forces between the ions and the polar water molecules. There are also ion-ion attractions between the solvated Na^+(aq) and Cl^-(aq); however, these forces are weaker than they were in NaCl(s) because the distance between the ions is greater in solution.

20.   There are three key points to address:

       General increase in normal boiling point as mass increases (as observed in Group 18; also observed in HCl vs. HBr vs. HI)

- All the elements in Group 18 are nonpolar. As such, the only intermolecular forces that can exist in a sample of pure liquid Group 18 element are induced dipole-induced dipole forces.

- Induced dipoles are larger in larger atoms/molecules than in smaller atoms/molecules. Here, we can use mass as a proxy for size.

- Since the induced dipoles are larger for the larger atoms, we see that the normal boiling point increases as size increases. (Ne < Ar < Kr < Xe _and_ HCl < HBr < HI)

   Higher normal boiling points for HX than for Group 18 element of same mass

- HX molecules are polar. As such, we also have to consider dipole-dipole IMF and dipole-induced dipole IMF.

- Dipole-dipole (and dipole-induced dipole) IMF are generally stronger than induced dipole-induced dipole IMF therefore it is not surprising that we see the normal boiling points of each HX as higher than the normal boiling points of the corresponding Group 18 element.

       Unusually high normal boiling point of HF

- HF is the only substance on the graph that can experience hydrogen bonding (since F is the only halogen small enough).

- Hydrogen bonding is generally a much stronger IMF than other dipole-dipole IMFs.

- This is why the normal boiling point for HF breaks the trend of boiling point increasing with mass.

21.

(a)
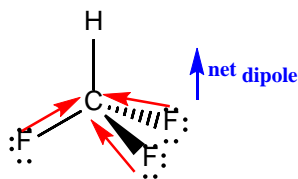
(b)   *see above*

A small C-H dipole pointing toward H is acceptable but not necessary.

(c) CHF_3 is the most polar of the haloforms because the C-F bonds are more polar than C-Cl, C-Br or C-I. This is due to the fact that the electronegativity difference between C and F is greatest (which is, in turn, due to fluorine being the most electronegative of the halogens).

(d)   The physical data in the table above indicate that the intermolecular forces between CHI_3 molecules are strongest (since it presumably has the highest boiling point and is the only haloform which is a solid at room temperature) while the intermolecular forces between CHF_3 molecules are weakest (since it has the lowest boiling point).

At first, this seems inconsistent with the answer to part (c) which implies that the dipole-dipole attractions must be strongest in CHF_3. These are not the only intermolecular forces that must be considered, however. Dipole-induced dipole and induced dipole-induced dipole attractions also contribute to the observed physical properties.

The induced dipoles will be greatest in CHI_3 because these are the largest molecules and contain the largest (and therefore most polarizable) halogen atoms. As such, the induced dipole-induced dipole forces will be strongest in CHI_3, and the physical data indicate that they contribute significantly to its physical characteristics.