Intermolecular Forces

Name: _____
Period: ___

(1) All atoms/molecules have _____ _____ _____.

Dipole interactions are seen between _____ molecules. Hydrogen bonds occur in molecules with hydrogen bonded to _____, _____, or _____.

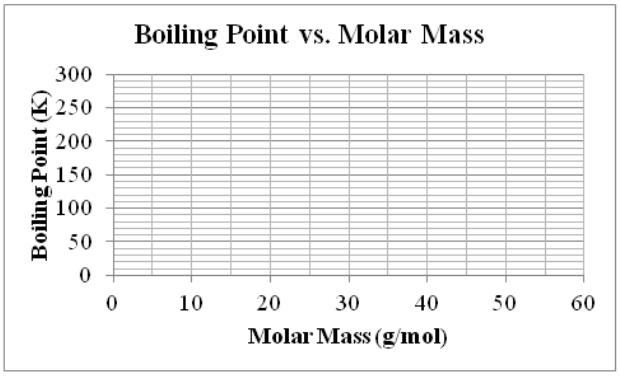
(2) Identify all of the intermolecular bonds present in each substance by placing an "X" in the appropriate column(s).

Substance	**Polar or non-polar?** *For diatomic molecules, calculate ΔEN. For polyatomic molecules, give the Lewis structure and classify the VSEPR shape for central atom(s).*	**London Dispersion Forces**	**Dipole Interactions**	**Hydrogen Bonding**
H_2O				
PCl_3				
C_3H_8				
CO_2				
NH_3				
He				
CH_3Cl				
SF_4				
I_2				
CH_4				
CH_3OH				
N_2				
CH_3NH_2				
CO				
N_2H_2				
C_2H_2				
XeF_4				
AsH_3				
SO				
HCN				

(3) Consider the elements neon and radon. Both are gases at room temperature (298 K) but neon has a boiling point at 27.1 K and radon has a boiling point of 211.5 K. What is the only type of intermolecular force observed between atoms of these elements? Why is radon observed to have a higher boiling point than neon?

(4) Consider the hydrocarbons methane (CH_4), ethane (C_2H_6), propane (C_3H_8), and butane (C_4H_{10}). In each of these molecules, what is the only type of intermolecular force observed? Calculate the molar mass of each compound. Plot the point on a graph of boiling point vs. molar mass. Explain the relationship between molar mass and boiling point.

Compound	Molar Mass	Boiling Point
CH_4		110 K
C_2H_6		185 K
C_3H_8		213 K
C_4H_{10}		274 K



(5) Consider the molecules hydrogen sulphide, H_2S, and Phosphane, PH_3.
(a) (i) Calculate the molar mass for each molecule and record in the table below.
(ii) Based on the result, compare the strength of the London Dispersion Forces for the two molecules.

(b) Draw the Lewis structure for each molecule in the table below. Classify the VSEPR shape of each molecule to confirm that both molecules are polar (have a dipole moment). List the types of intermolecular forces for each molecule.
(c) Why is hydrogen sulphide observed to have a higher boiling point than phosphane?

Molecule and molar mass	Lewis Structure and VSEPR shape	Intermolecular Forces	Dipole Moment	Boiling Point
H_2S _____	_____		0.97 D	213 K
PH_3 _____	_____		0.85 D	186 K

(6) (a) Dipole- induced dipole interactions are present between a _____ molecule and a _____ molecule. The strength of the dipole-induced dipole interactions increases with the permanent _____ moment of the polar molecule and the _____ of the non-polar molecule.

(b) Consider the molecules HCl and HI. Both HCl and HI are polar molecules. Calculate ΔEN for each molecule. How does the dipole moment relate to the value of ΔEN?

Molecule	ΔEN	Dipole Moment
HCl		1.05
HI		0.38

(c) Consider the molecules H_2 and F_2. Both of the molecules are non-polar. Which of the molecules would be the most polarizable? Explain.

(d) What pairing of one polar molecule (HCl or HI) and one non-polar molecule (H_2 or F_2) would produce the strongest dipole-induced dipole interaction?

(7) (a) Ion-dipole interactions are a result of the _____ attraction between an ion and a _____ molecule. The strength of these interactions is higher for ions that have a _____ ionic radius and for ions that have a _____ charge.

(b) Consider the following ions:

Ion	Ionic Radius (pm)
Mg^{2+}	86
Ca^{2+}	100
Sr^{2+}	132

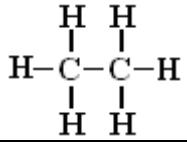
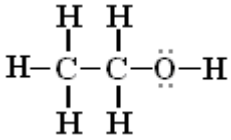
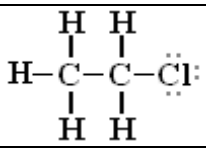
Rank the ions in order of their relative strength of attraction with water molecules. Explain.

(c) Consider the following ions:

Ion	Ionic Radius (pm)
Li^{+}	76
Zn^{2+}	74
Sc^{3+}	75

Rank the ions in order of their relative strength of attraction with water molecules. Explain.

(8) Complete the following table. Explain the relative boiling points/heats of vapourization in terms of intermolecular forces. Which substance would be the most soluble in water?

Molecule	Lewis Structure	Intermolecular Forces	Boiling Point	H_{vap}
Ethane			–89 ºC	14 kJ/mol
Ethanol			78 ºC	39 kJ/mol
Chloroethane			12 ºC	26 kJ/mol

Answers:

(1) All atoms/molecules have London Dispersion Forces. Dipole interactions are seen between polar molecules. Hydrogen bonds occur in molecules with hydrogen bonded to fluorine, oxygen, or nitrogen.

(2)

Substance	Polar or non-polar?	LDF	Dipole	H-Bonds
H_2O	polar, bent	X	X	X
PCl_3	polar, trigonal pyramidal	X	X	
C_3H_8	non-polar, tetrahedral	X		
CO_2	non-polar, linear	X		
NH_3	polar, trigonal pyramidal	X	X	X
He	non-polar	X		
CH_3Cl	polar, asymmetric tetrahedral	X	X	
SF_4	polar, irregular tetrahedral	X	X	
I_2	non-polar, $\Delta EN = 0$	X		
CH_4	non-polar, tetrahedral	X		
CH_3OH	polar	X	X	X
N_2	non-polar, $\Delta EN = 0$	X		
CH_3NH_2	polar	X	X	X
CO	$\Delta EN = 1.0$	X	X	
N_2H_2	polar	X	X	X
C_2H_2	non-polar	X		
XeF_4	non-polar, square planar	X		
AsH_3	polar, trigonal pyramidal	X	X	
SO	polar, $\Delta EN = 1.0$	X	X	
HCN	polar, asymmetric linear	X	X	

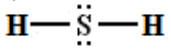
(3) Both atoms have only London dispersion forces. Radon (n = 6) has a larger atomic radius than neon (n = 2) . The larger the atom, the more polarizable it is and the stronger the London dispersion forces. Since the London dispersion forces are stronger for radon than for neon, radon has a higher boiling point than neon.

(4) The molecules have only London dispersion forces. The higher the molar mass, the higher the boiling point. The larger a molecule is (the more electrons a molecule has), the stronger the London dispersion forces and the higher the boiling point.

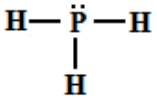
(5) (a) (i) H_2S = 34.076 g/mol PH_3 = 33.994 g/mol

(ii) The molecules have comparable molar masses and would have London dispersion forces of very similar strength.

(b)



bent, polar, London dispersion forces, dipole interactions



trigonal pyramidal, polar, London dispersion forces, dipole interactions

(c) H_2S is observed to have a higher dipole moment than PH_3, as a result the H_2S has stronger dipole interactions than PH_3 and H_2S will have a higher boiling point than PH_3.

(6) (a) Dipole- induced dipole interactions are present between a polar molecule and a non-polar molecule. The strength of the dipole-induced dipole interactions increases with the permanent dipole moment of the polar molecule and the polarizability of the non-polar molecule.

(b) HCl: $\Delta EN = 0.90$, HI: $\Delta EN = 0.40$. The greater the ΔEN, the larger the dipole moment (i.e. the more polar the molecule).

(c) F_2 would be more polarizable than H_2 because F_2 is a larger molecule and has more electrons than H_2.

(d) The strongest dipole-induced dipole interaction would occur for HCl with F_2.

(7) (a) Ion-dipole interactions are a result of the electrostatic attraction between an ion and a polar molecule. The strength of these interactions is higher for ions that have a smaller ionic radius and for ions that have a higher charge.

(b) $Sr^{2+} < Ca^{2+} < Mg^{2+}$ All the ions have the same charge, so the strength of attraction with water is greater for smaller ions.

(c) $Li^{+} < Zn^{2+} < Sc^{3+}$ All the ions have approximately the same size, so the strength of attraction with water is greater for the higher charge.

(8) Ethane: London dispersion forces; Ethanol: London dispersion forces, dipole interactions, and hydrogen bonding, Chlorothane: London dispersion forces and dipole interactions. Ethane has the lowest boiling point/heat of vapourization because the molecules are held together only by London dispersion forces, the weakest type of intermolecular bonds, which are easily broken. Chloroethane has an intermediate boiling point/heat of vapourization as the molecules are held together by both London dispersion forces and dipole interactions. The dipole interactions result in stronger intermolecular bonds. Ethanol has the highest boiling point/heat of vapourization as the molecules participate in hydrogen bonding, the strongest type on intermolecular bond.

Ethanol would be the most soluble in water since has the capacity to form hydrogen bonds with water molecules. None of the other molecules would be able to form hydrogen bonds with water, so they would be less soluble/insoluble.