CHAPTER 11 – Liquids and Intermolecular Forces

Section 11.1 – A Molecular Comparison of Gases, Liquids, and Solids

(a) According to the kinetic molecular theory, the attractive forces between gas particles for an ideal gas are ___________________________.

(b) Compare the properties of gases, liquids, and solids by filling in the table below.

	Gas	Liquid	Solid
Does it have a definite volume?			
Does it have a definite shape?			
Does it expand to fill its container?			
Is it compressible?			
Do its particles flow readily?			
Is diffusion fast or slow?			
Is its density relatively high or low?			

(c) Because the particles in a solid or liquid are fairly close together compared with those of a gas, we often refer to solids and liquids as ______________________________________ phases

(d) Based on the information in Figure 11.2, describe how the intermolecular attractive forces appear to change as one moves from chlorine to bromine to iodine.

(e) The kinetic energies of particles are usually controlled by changing the ____________________ of a substance.

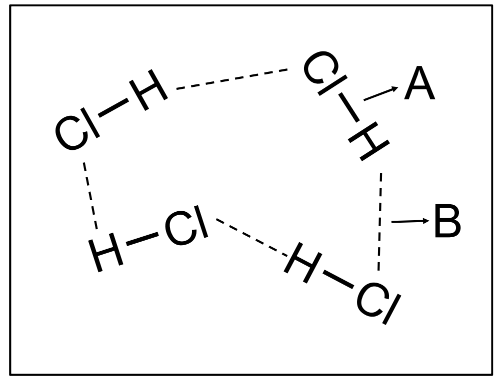
(f) Describe what happens to the attractive forces between water molecules as water vapor undergoes condensation, turning into drops of liquid water.

(g) Explain how increasing the pressure on a sample of gas can affect the attractive forces between particles.

Section 11.2 – Intermolecular Forces

(a) How does the strength of intermolecular forces (in general) compare with the strength of ionic, metallic, or covalent bonds?

(b) Referring to the diagram at right, discuss the difference in strength between the attractive forces labeled as A and B.



Enthalpy of Vaporization for HCl	16 kJ/mol
Bond Enthalpy for H–Cl	431 kJ/mol

(c) Explain how the data in the table above supports your answer to part (b).

TABLE 11.2 • Melting and Boiling Points of Representative Substances

Force Holding Particles Together	Substance	Melting Point (K)	Boiling Point (K)
Chemical bonds			
Covalent bonds	Diamond (C)	3800	4300
Metallic bonds	Beryllium (Be)	1560	2742
Ionic bonds	Lithium fluoride (LiF)	1118	1949
Intermolecular forces			
Dispersion force	Nitrogen (N_2)	63	77
Dipole–dipole force	Hydrogen chloride (HCl)	158	188
Hydrogen bonding force	Hydrogen fluoride (HF)	190	293

(d) What conclusions can be made about the relative strengths of attractive forces from the information in Table 11.2?

(e) Three types of intermolecular attractive forces exist between neutral molecules. Name them.

(f) Electrostatic interactions can be described by Coulomb's law. Give a brief summary of how electrostatic attractions are affected by the magnitude of the charges and the interparticle distance.

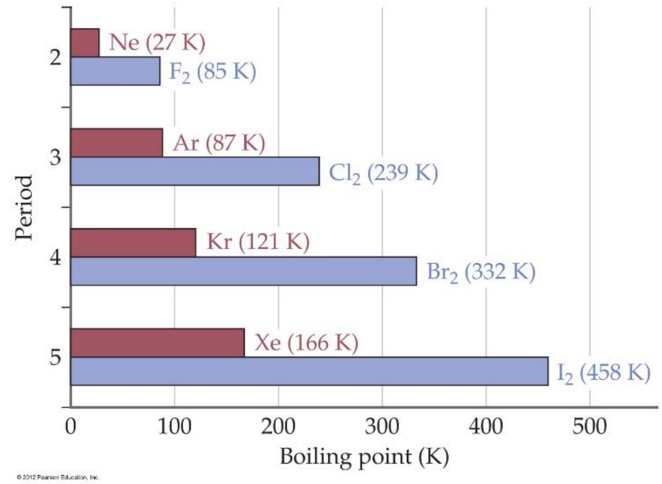
(g) Based on your answer to part (f), give two specific reasons why intermolecular forces are weaker than ionic bonds.

(h) In 1930, Fritz London proposed that the motion of electrons in an atom or molecule can create an ________________________, or momentary dipole moment.

(i) Describe what can happen in an atom or a molecule when the distribution of the electrons can be arranged asymmetrically, even for just an instant.

(j) Define polarizability and explain how this property is related to the relative strength of dispersion forces.

(k) Use your answer to part (j) to explain the boiling point data shown in Figure 11.5



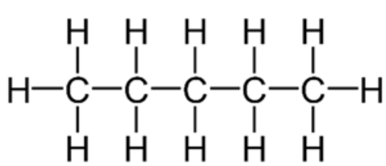
****Note: The polarizability of a molecule and the dispersion forces generally increase with the <u>size</u> of the atom or molecule. When you compare two substances on the basis of their London dispersion forces, you should always talk about the size of the electron cloud and the polarizability. Do NOT explain a BP trend as follows.**

"X has a higher boiling point than Y because it has a greater mass."

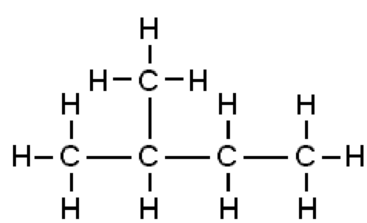
That explanation would be incorrect. It's not about the mass. It's about the polarizability of the electron cloud.

(l) Dispersion forces increase with contact area between molecules. The shape of a molecule affects its surface area and how much it comes into contact with its neighboring molecules. Which of the following molecules has the lowest BP? the highest BP?

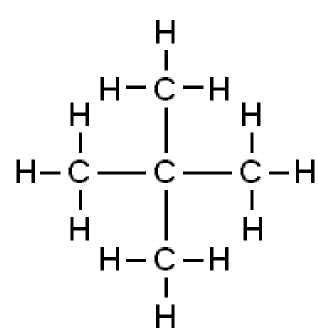
(A)



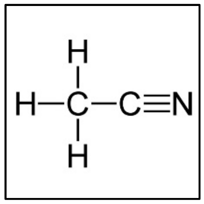
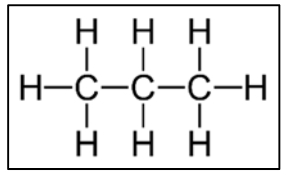
(B)



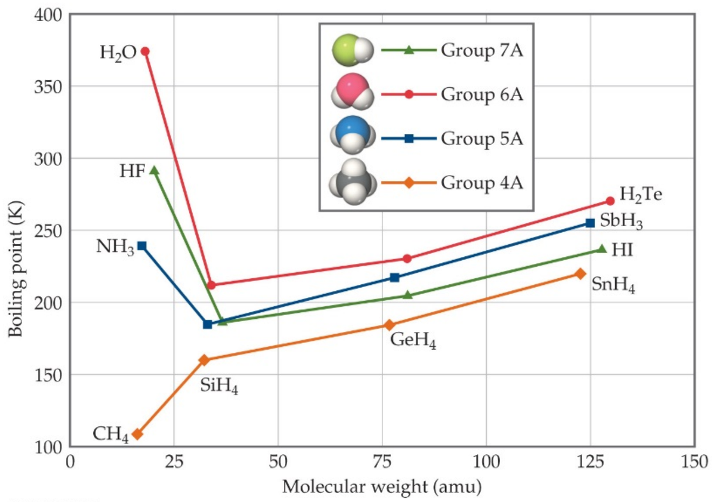
(C)



(m) Dipole-dipole forces occur in polar molecules. Draw the structural formula of two $CHCl_3$ molecules. Label the positive and negative ends of each molecule. Use a dashed line to show the dipole-dipole force between them.

(n) Identify which of the substances shown above has a higher boiling point. Justify your answer based on the structural features of each molecule and the type of intermolecular forces that each one experiences.

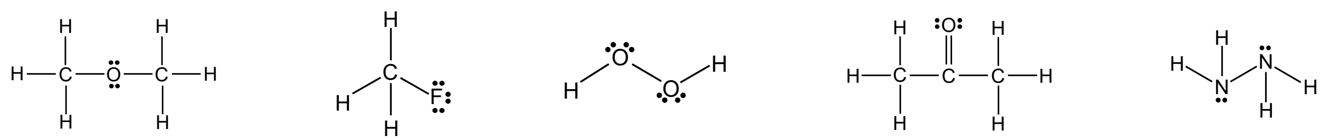


(o) Explain the BP trend in the Group 4A (Group 14) hydrides.

(p) Which three hydride molecules in the graph have unusually high boiling points compared to the others that are in the same series? Why do these molecule have such high boiling points?

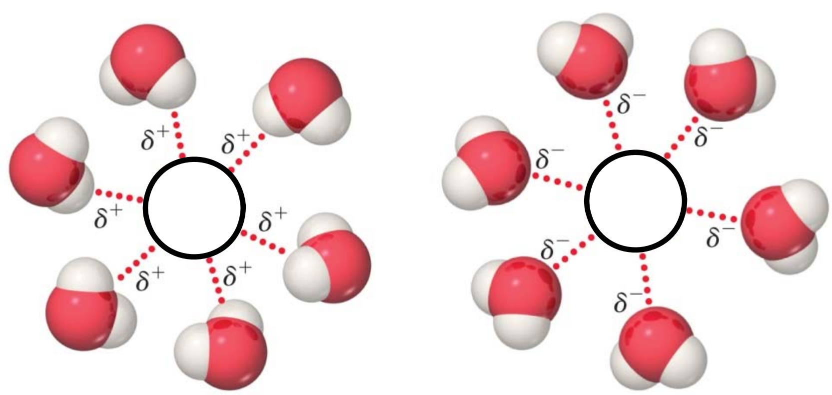
(q) Draw two ammonia molecules. Using a dashed line, draw a hydrogen bond between them. Make sure you know the difference between a hydrogen bond and a covalent bond. Refer to Figure 11.10.

(r) Circle the molecules below that are able to participate in hydrogen bonding.



(s) How does hydrogen bonding explain why water expands when it freezes?

(t) An ion-dipole force exists between an ion and a polar molecule like water. Label each ion below with its charge.



(u) All atoms and molecules experience ________________________ forces. Only polar molecules

experience ____________________________ forces. And only a molecule that has a hydrogen

directly bonded to an N, O, or F will experience ____.

If an ionic substance is dissolved in a polar solvent, then _______________________________

forces will be present.

(v) Match the following six substances with their boiling point. Justify your answer.

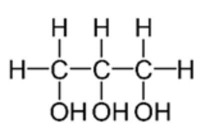
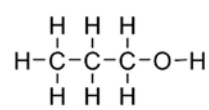
The boiling point values are –161°C, –42°C, –3°C, 98°C, 211°C, and 2850°C

1-propanol, $CH_3CH_2CH_2OH$	propane, $CH_3CH_2CH_3$
1-fluoropropane, $CH_3CH_2CH_2F$	1,3-propanediol, $HOCH_2CH_2CH_2OH$
methane, CH_4	calcium oxide, CaO

Section 11.3 – Select Properties of Liquids

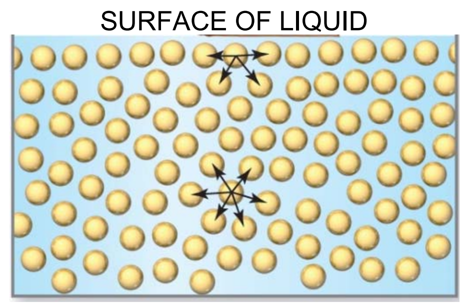
(a) Define viscosity. ____

(b) Explain why decane ($C_{10}H_{22}$) has a greater viscosity than hexane (C_6H_{14}).

(c) Explain why glycerol  has a greater viscosity than propanol 

(d) Explain how the viscosity of a liquid is affected by increasing the temperature.

(e) Use the diagram at right to explain why certain liquids experience surface tension.



(f) What are the attractive forces experienced by these liquids, both of which have high surface tension?

(i) water ______________________________ (ii) mercury ___________________________________

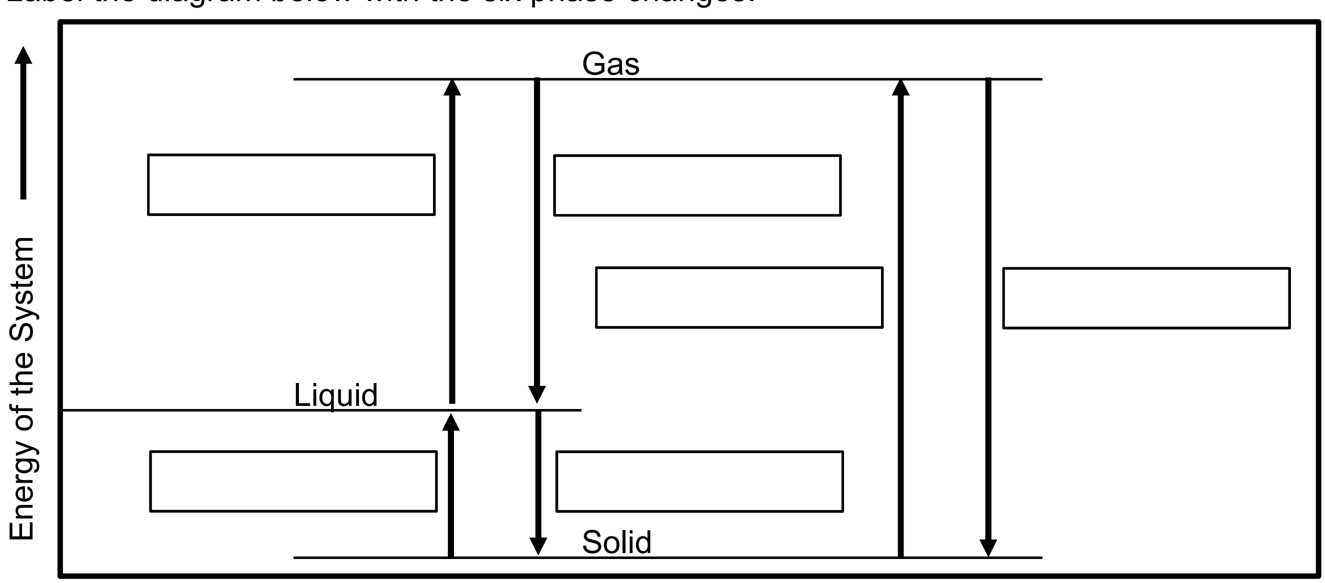
(g) See figure at right. The meniscus of water looks like this because the adhesive forces (H_2O – glass) are (less more) than the cohesive forces between H_2O molecules.

The meniscus of mercury looks like this because the adhesive forces (Hg – glass) are (less more) than the cohesive forces between Hg atoms.

(h) Explain how capillary action works.

Section 11.4 – Phase Changes

(a) Label the diagram below with the six phase changes.

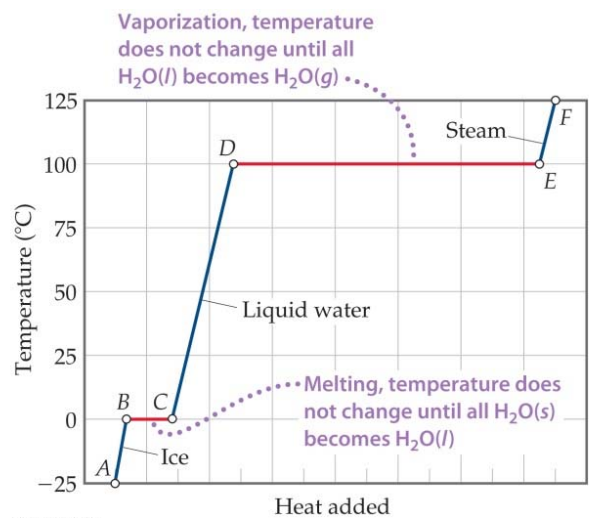


(b) Which phase changes are endothermic? Which are endothermic?

(c) Define enthalpy of fusion and enthalpy of vaporization.

(d) Explain why ΔH_{vap} tends to be larger than ΔH_{fus}.

(e) Identify the segments of this heating curve that would involve the equation
$q = m \times C \times \Delta T$



(f) Calculate how many kilojoules of heat is needed to convert 9.0 grams of ice, $H_2O(s)$ at –10°C into steam, $H_2O(g)$, at 110°C. The specific heat values for ice, liquid water, and steam are 2.03 J/g-K, 4.18 J/g-K, and 1.84 J/g-K respectively. The heat of fusion for H_2O is 6.01 kJ/mol and the heat of vaporization for H_2O is 40.67 kJ/mol.

(g) Explain the phenomenon known as *supercooling*.

(h) Explain why a substance cannot become liquefied at a temperature above the critical temperature.

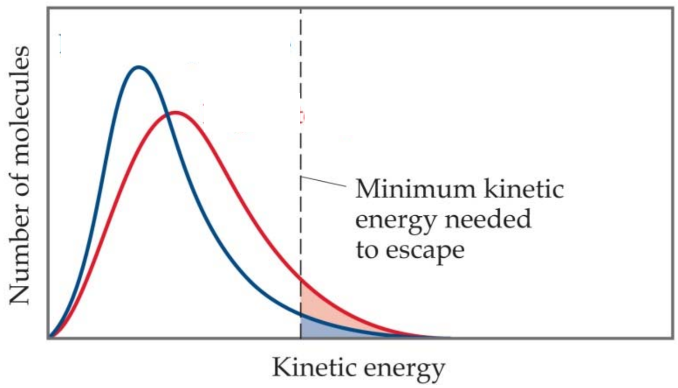
Section 11.5 – Vapor Pressure

(a) Suppose that a sample of liquid that is placed in an evacuated, closed container. Describe what happens to the liquid over time, and explain why the vapor pressure will eventually reach a constant value.

(b) What is the dynamic equilibrium that occurs in the situation described in part (a)?

(c) Define the term volatile as it applies to a liquid. What does volatility suggest about the strength of a liquid's intermolecular forces?

(d) Which curve on this diagram represents a higher temperature?



(e) Use the diagram above to explain how the vapor pressure of a liquid is affected by an increase in temperature.

(f) A liquid starts to boil when its vapor pressure is equal to ______________________________

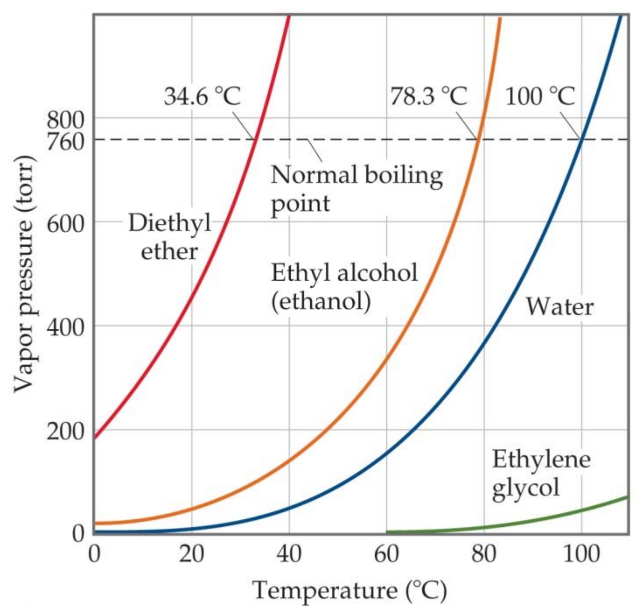
(g) The normal boiling point is defined at what pressure? _______________

(h) Describe how the boiling point of water is affected by the following situations.

(i) high altitudes ____

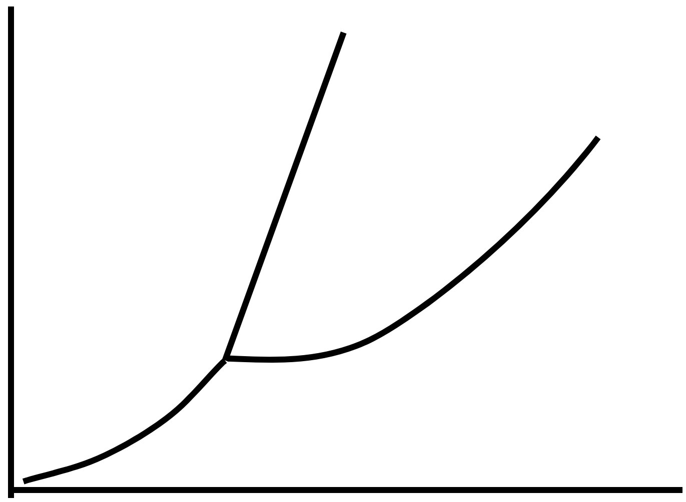
(ii) a pressure cooker ____

(i) Which of these liquids has the weakest intermolecular forces? Which liquid has the strongest IMFs?



Section 11.6 – Phase Diagrams

(a) A phase diagram is a graphic way to summarize the conditions under which equilibria exist between the different states of matter. On the generic phase diagram below, label the x-axis and y-axis, identify the solid, liquid, and gas regions, and label the triple point and critical point.



Section 11.7 – Liquid Crystals

This section presents information that is interesting, but not included in the AP Chemistry curriculum and will not be covered on the AP Exam.