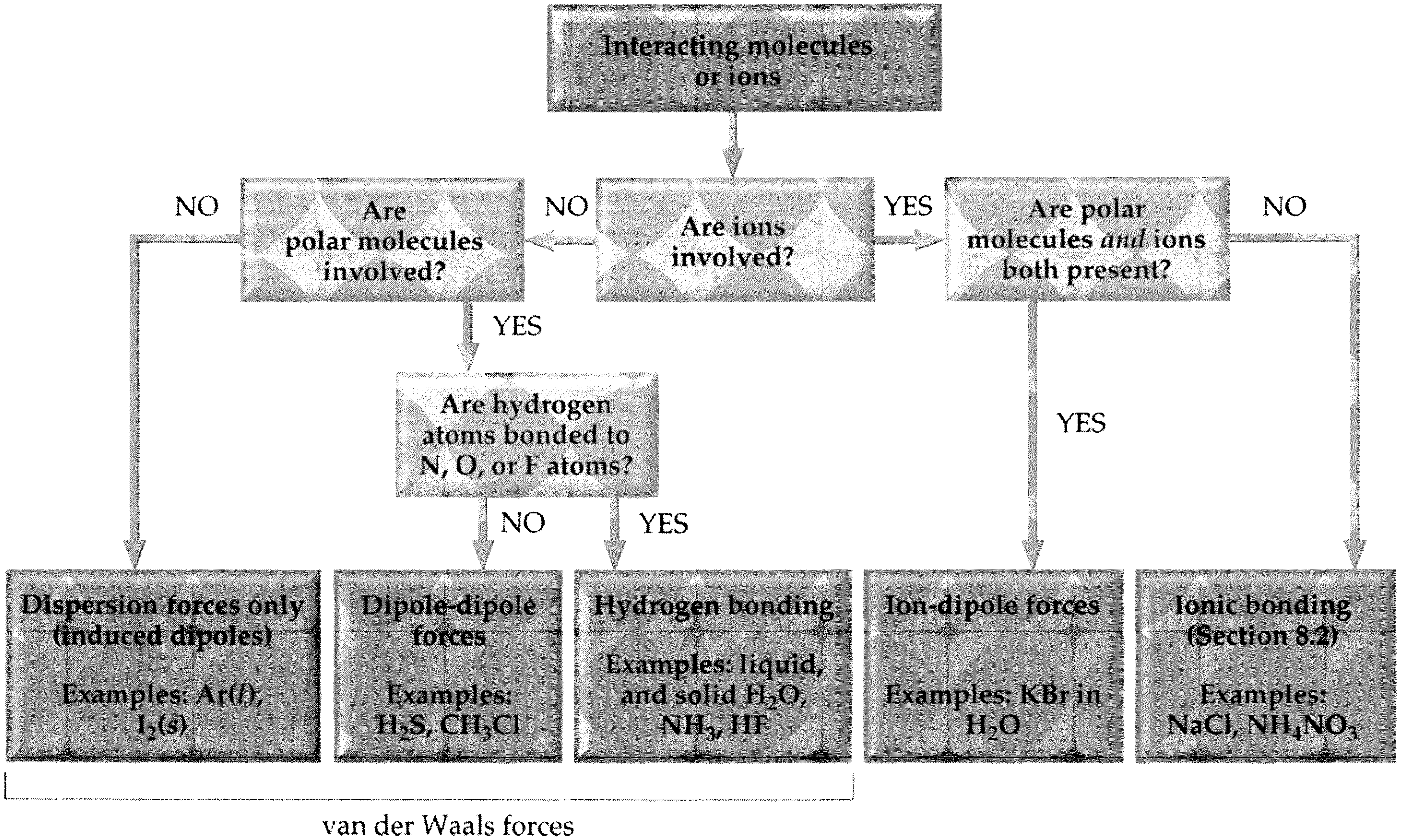
Summarizing Intermolecular Forces



Intermolecular Forces

The Periodic Table of the Elements (with Electronegativities)

- Alkali metals
- Alkaline earth metals
- Transition metals
- Lanthanides
- Actinides
- Other metals
- Metalloids (semi-metal)
- Nonmetals
- Halogens
- Noble gases

Element name → Mercury
Atomic # → 80
Symbol → Hg
Avg. Mass → 200.59
Electronegativity → 1.9

1	2		3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
Hydrogen 1 H 1.01 2.1																		Helium 2 He 4.00 —
Lithium 3 Li 6.94 1.0	Beryllium 4 Be 9.01 1.5												Boron 5 B 10.81 2.0	Carbon 6 C 12.01 2.5	Nitrogen 7 N 14.01 3.0	Oxygen 8 O 16.00 3.5	Fluorine 9 F 19.00 4.0	Neon 10 Ne 20.18 —
Sodium 11 Na 22.99 0.9	Magnesium 12 Mg 24.31 1.2												Aluminum 13 Al 26.98 1.5	Silicon 14 Si 28.09 1.8	Phosphorus 15 P 30.97 2.1	Sulfur 16 S 32.07 2.5	Chlorine 17 Cl 35.45 3.0	Argon 18 Ar 39.95 —
Potassium 19 K 39.10 0.8	Calcium 20 Ca 40.08 1.0		Scandium 21 Sc 44.96 1.3	Titanium 22 Ti 47.88 1.5	Vanadium 23 V 50.94 1.6	Chromium 24 Cr 52.00 1.6	Manganese 25 Mn 54.94 1.5	Iron 26 Fe 55.85 1.8	Cobalt 27 Co 58.93 1.8	Nickel 28 Ni 58.69 1.8	Copper 29 Cu 63.55 1.9	Zinc 30 Zn 65.39 1.6	Gallium 31 Ga 69.72 1.6	Germanium 32 Ge 72.61 1.8	Arsenic 33 As 74.92 2.0	Selenium 34 Se 78.96 2.4	Bromine 35 Br 79.90 2.8	Krypton 36 Kr 83.80 3.0
Rubidium 37 Rb 85.47 0.8	Strontium 38 Sr 87.62 1.0		Yttrium 39 Y 88.91 1.2	Zirconium 40 Zr 91.22 1.4	Niobium 41 Nb 92.91 1.6	Molybdenum 42 Mo 95.94 1.8	Technetium 43 Tc (98) 1.9	Ruthenium 44 Ru 101.07 2.2	Rhodium 45 Rh 102.91 2.2	Palladium 46 Pd 106.42 2.2	Silver 47 Ag 107.87 1.9	Cadmium 48 Cd 112.41 1.7	Indium 49 In 114.82 1.7	Tin 50 Sn 118.71 1.8	Antimony 51 Sb 121.76 1.9	Tellurium 52 Te 127.60 2.1	Iodine 53 I 126.90 2.5	Xenon 54 Xe 131.29 2.6
Cesium 55 Cs 132.91 0.7	Barium 56 Ba 137.33 0.9	57-70 *	Lutetium 71 Lu 174.97 1.1	Hafnium 72 Hf 178.49 1.3	Tantalum 73 Ta 180.95 1.5	Tungsten 74 W 183.84 1.7	Rhenium 75 Re 186.21 1.9	Osmium 76 Os 190.23 2.2	Iridium 77 Ir 192.22 2.2	Platinum 78 Pt 195.08 2.2	Gold 79 Au 196.97 2.4	Mercury 80 Hg 200.59 1.9	Thallium 81 Tl 204.38 1.8	Lead 82 Pb 207.20 1.8	Bismuth 83 Bi 208.98 1.9	Polonium 84 Po (209) 2.0	Astatine 85 At (210) 2.2	Radon 86 Rn (222) 2.4
Francium 87 Fr (223) 0.7	Radium 88 Ra (226) 0.9	89-102 **	Lawrencium 103 Lr (262) —	Rutherfordium 104 Rf (267) —	Dubnium 105 Db (268) —	Seaborgium 106 Sg (271) —	Bohrium 107 Bh (272) —	Hassium 108 Hs (270) —	Meitnerium 109 Mt (276) —	Darmstadtium 110 Ds (281) —	Roentgenium 111 Rg (280) —	Copernicium 112 Cn (285) —	Ununtrium 113 Uut (284) —	Ununquadium 114 Uuq (289) —	Ununpentium 115 Uup (288) —	Ununhexium 116 Uuh (293) —	Ununseptium 117 Uus (294?) —	Ununoctium 118 Uuo (294) —

*lanthanides	Lanthanum 57 La 138.91 1.1	Cerium 58 Ce 140.12 1.1	Praseodymium 59 Pr 140.91 1.1	Neodymium 60 Nd 144.24 1.1	Promethium 61 Pm (145) 1.1	Samarium 62 Sm 150.36 1.2	Europium 63 Eu 151.97 1.1	Gadolinium 64 Gd 157.25 1.2	Terbium 65 Tb 158.93 1.1	Dysprosium 66 Dy 162.50 1.2	Holmium 67 Ho 164.93 1.2	Erbium 68 Er 167.26 1.2	Thulium 69 Tm 168.93 1.3	Ytterbium 70 Yb 173.04 1.1
**actinides	Actinium 89 Ac (227) 1.1	Thorium 90 Th 232.04 1.3	Protactinium 91 Pa 231.04 1.5	Uranium 92 U 238.03 1.4	Neptunium 93 Np (237) 1.4	Plutonium 94 Pu (244) 1.3	Americium 95 Am (243) 1.3	Curium 96 Cm (247) 1.3	Berkelium 97 Bk (247) 1.3	Californium 98 Cf (251) 1.3	Einsteinium 99 Es (252) 1.3	Fermium 100 Fm (257) 1.3	Mendelevium 101 Md (258) 1.3	Nobelium 102 No (259) 1.3

Endothermic and Exothermic Reactions

Objective:

Investigate the difference between endothermic and exothermic reactions.

Chemicals:

- Tap water
- Calcium chloride chips
- Ammonium nitrate granules

Materials:

- Spark unit
- Temperature sensor
- Stainless steel temperature probe
- Wire test tube rack
- 4 large test tubes

Procedure:

1. Obtain all materials.
2. Start the Spark. Plug the stainless steel probe into the temperature sensor. Attach the sensor into the Spark.
3. Once a temperature is displayed on the screen, select the temperature. The temperature will be highlighted in orange. Select show. A graphical display should appear on the Spark. You are now ready to collect temperature data.
4. Half fill two test tubes with tap water. Place into the test tube rack.
5. Carefully using a scupula, add calcium chloride chips to an empty test tube until it is 1/8 full. Place the test tube into the rack.
6. Carefully using a scupula, add ammonium nitrate granules to an empty test tube until it is 1/8 full. Rack the test tube.
7. Place the temperature probe into one of the test tubes containing solid. Press one of the silver buttons on the Spark to begin collecting temperature data.
8. Wait for the temperature to stabilize then carefully add the water from one of the test tubes to your solid. Use the probe to stir the solid until once again the temperature stabilizes.
9. Stop collecting data on the Spark.
10. Press the blue graph icon in the bottom left of the screen. A menu should appear on the right of the screen. Select the icon at the top right of the menu. Your graph should enlarge to fit the whole screen. Then select the sigma icon (it looks like this Σ). A graph statistics window should appear. Select minimum and maximum from the options and OK.
11. Sketch the graph including minimum and maximum points on the graph paper provided.
12. Remove the probe form the solution, rinse and dry it. Repeat steps 7 to 11 for the second solid.
13. Turn off the Spark and rinse all solutions in the sink with lots of water. Rinse and dry the probe and put all equipment away.

W3 #3

CH 30S

Change of State

1. Use the following chart and answer the questions that follow

Substance	Melting Point	Boiling Point
Water	0 °C	100 °C
Para -dichloro benzene	53 °C	174 °C
Napthalene	80 °C	218 °C
Methyl Alcohol	-98 °C	165 °C
Mercury	-39 °C	357 °C

a) What will the state of the following substances be under the following conditions? Solid (s), Liquid (L) or Gas (G) ?

Substance	Conditions				
	0°C	Room Temp	55 °C	100 °C	400 °C
Water					
Napthalene					
Para -dichloro benzene					
Methyl Alcohol					
Mercury					

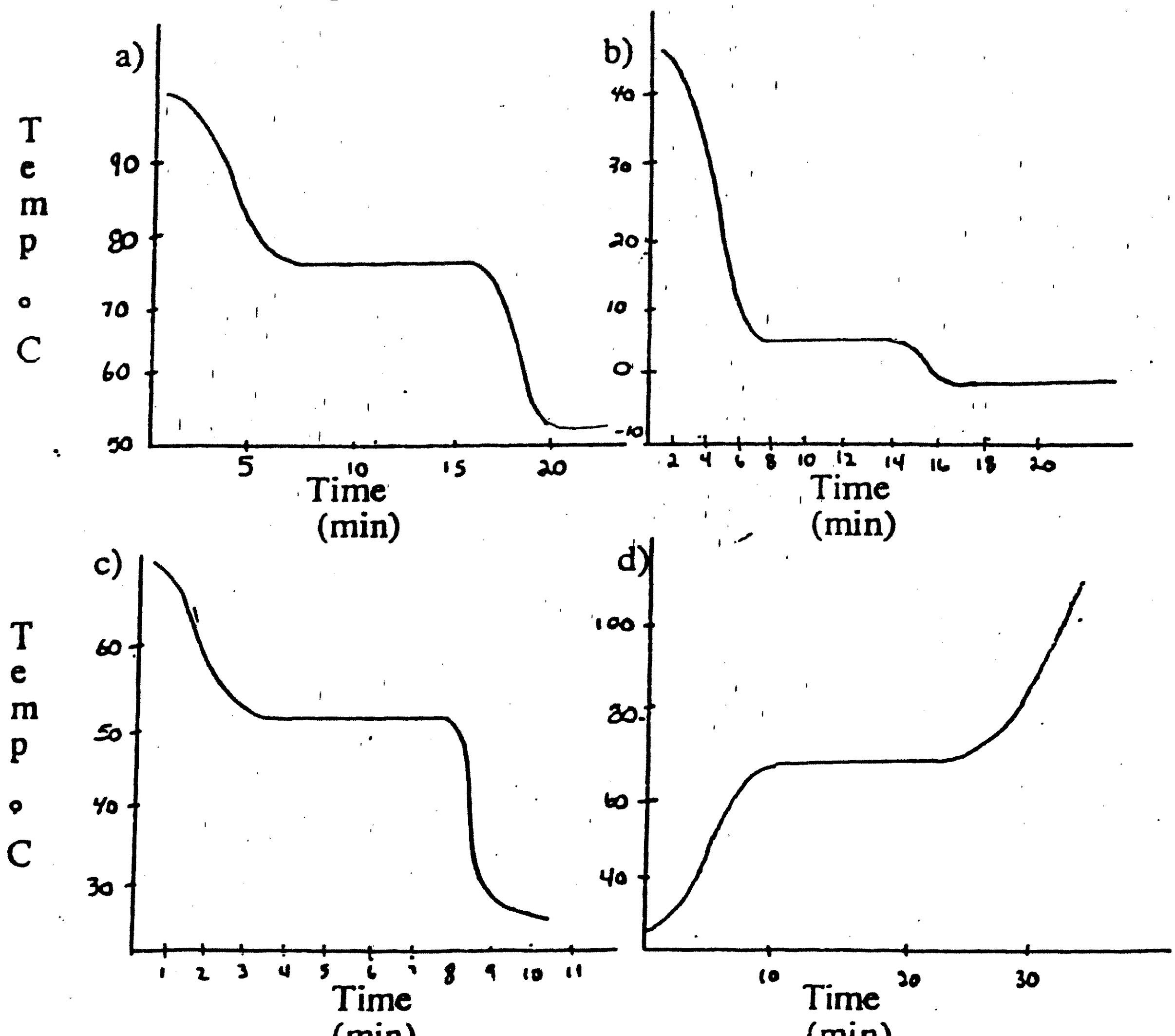
b) In what order would Ice, Para and Naptha melt?

c) In what order would Water, Para and Naptha freeze?

2. Which of the following has the highest Melting point?
   a) 0.02 g of Naptha
   b) 2.0 g of Naptha
   c) 20 g of Naptha

3. What would be the freezing point of Mercury and methyl Alcohol?

4. What is the **Freezing Point** of the following substances?

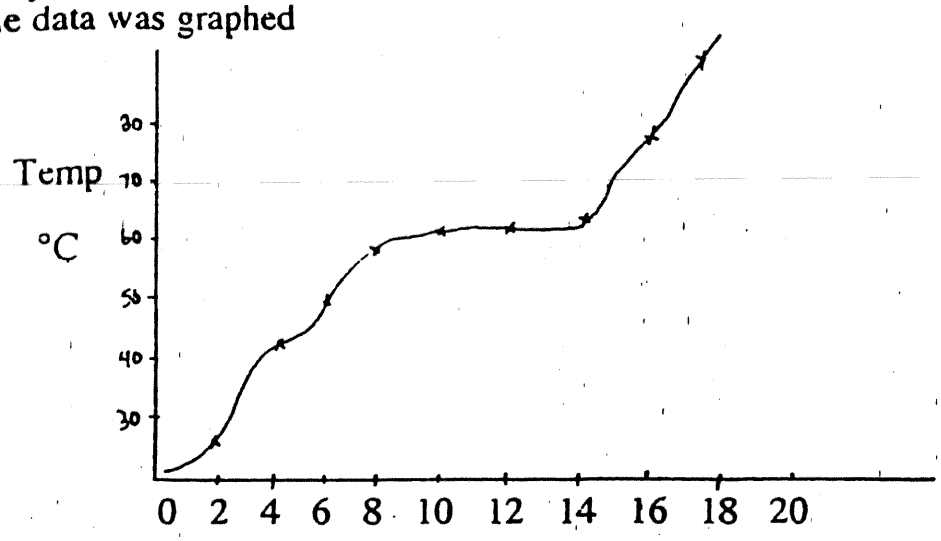


a) Which of the above could be Para?

b) Why do some graphs have two plateaus?

c) Indicate on each graph where solids, solids and liquids and liquids co-exist.

5. A student melted a material using one burner , 130 ml of water as a water bath, 5.00 grams of material and took temperature readings every 0.5 minute interval.
The data was graphed



Indicate on the graph what you would expect to happen if:

a) two burners were used?

b) if 10.00 g of material were used?

c) if 1.00 g of material was used?

d) if 5.00 grams of the same material was cooled from 70 °C to 40 °C?

6. Which of the following variables affect the Freezing Point of a substance?

Variable	Effect on the Freezing Point
a) Heat of the burner	
b) Time of cooling	
c) Amount of Material	
d) Type of Thermometer used	
e) Type of material	
f) Strength of intermolecular bonds	

PROBLEMS

6. Convert the following temperatures from Celsius to kelvin.
   a. 65° b. 16° c. 48° d. −36° e. −73°
7. Convert the following temperatures from kelvin to Celsius.
   a. 86 b. 191 c. 533 d. 321 e. 894
8. Convert the following temperatures from Celsius to kelvin.
   a. 23° b. 58° c. −90° d. 18° e. 25°
9. Convert the following temperatures from kelvin to Celsius.
   a. 872 b. 690 c. 384 d. 20 e. 60
10. At 25°C, which of the following gas molecules move fastest?
   a. N_2 b. F_2 c. CO_2 d. O_2

6. a. 338 K b. 289 K c. 321 K d. 237 K e. 200 K
7. a. −187°C b. −82°C c. 260°C d. 48°C e. 621°C

15:5 STATES OF MATTER

Matter exists in four states—solid, liquid, gas, and plasma. Thus far, our discussion of the kinetic theory has been limited to gases. However, kinetic theory can also be used to explain the behavior of solids and liquids. Plasmas are treated as a special case.

Four states of matter: solid, liquid, gas, plasma.

Gas particles are independent of each other and move in a straight line. Change of direction occurs only when one particle collides with another, or when a particle collides with the walls of the container. Gas particles, then, travel in a completely random manner. Since they travel until they collide with a neighbor or with the walls of their container, gases assume the shape and volume of their container.

Gas particles travel in random paths.

Gases assume the shape and volume of their container.

The particles of a liquid have what appears to be a vibratory type of motion. Actually, they are traveling a straight-line path between collisions with near neighbors. The point about which the seeming vibration occurs often shifts as one particle slips past another. These differences in the amount of space between particles allow the particles to change their relative positions continually. Thus, liquids, although they have a definite volume, assume the shape of their container.

Liquid particles travel in straight-line paths between collisions, but appear to vibrate about moving points.

Liquids have definite volume but assume the shape of their container.

In solids, a particle occupies a relatively fixed position in relation to the surrounding particles. A particle of a solid appears to vibrate about a fixed point. Again, the particle is actually traveling a straight-line path between collisions with very near neighbors. For example, a molecule of oxygen gas at 25°C travels an average distance equal to 314 times its own diameter before colliding with another molecule. In a solid, however, the particles are closely packed and travel a distance equal to only a fraction of their diameters before colliding. Unlike liquids, solids have their particles arranged in a definite pattern. Solids, therefore, have both a definite shape and a definite volume.

Solid particles appear to vibrate about fixed points.

Solid particles are arranged in a definite pattern.

Solids have both definite shape and definite volume.

The physical state of a substance at room temperature and standard atmospheric pressure depends mostly on the bonding in the substance.

8

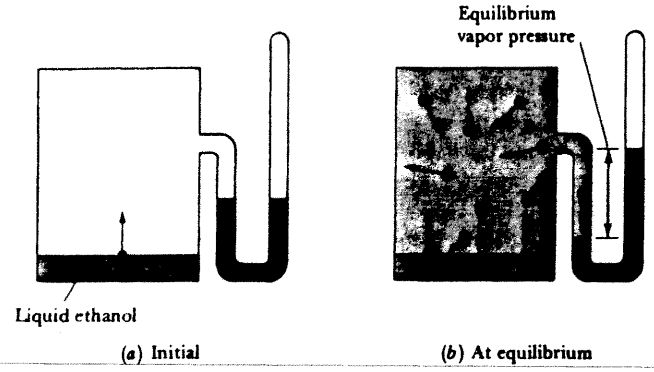


Figure 11.16 Illustration of the equilibrium vapor pressure over liquid ethanol. In (*a*), we imagine that no ethanol molecules exist in the gas phase; there is zero pressure in the cell. In (*b*), the rate at which molecules of ethanol leave the surface equals the rate at which gas molecules pass into the liquid phase. Thus, the rates of condensation and of vaporization are equal. This produces a stable vapor pressure that does not change with time, as long as temperature remains constant.

Explaining Vapor Pressure on the Molecular Level

The molecules of a liquid move at various speeds. Figure 11.17 shows the distribution of kinetic energies of the particles at the surface of a liquid at a particular temperature. The distribution curve is like those shown earlier for gases (Figure 10.13). At any instant, some of the molecules on the surface of the liquid possess sufficient energy to escape from the attractive forces of their neighbors. The weaker the attractive forces, the larger the number of molecules that are able to escape into the gas phase, and hence the higher the vapor pressure.

The movement of molecules from the liquid to the gas phase goes on continuously. However, as the number of gas-phase molecules increases, the probability increases that a molecule in the gas phase will strike the liquid surface and stick there [Figure 11.16(*b*)]. Eventually, the number of molecules returning to the liquid exactly equals the number escaping from it. The number of molecules in the gas phase then reaches a steady value, and the pressure of the vapor at this stage becomes constant.

The condition in which two opposing processes are occurring simultaneously at equal rates is called a **dynamic equilibrium**. A liquid and its vapor are in equilibrium when evaporation and condensation occur at equal rates. The observer may conclude that nothing is occurring during an equilibrium, because there is no net change in the system. In fact, a great deal is happening; molecules continuously pass from the liquid state to the gas state and from the gas state to the liquid state. All equilibria between different states of matter possess this **dynamic character**. ***The vapor pressure of a liquid is the pressure exerted by its vapor when the liquid and vapor states are in dynamic equilibrium.***

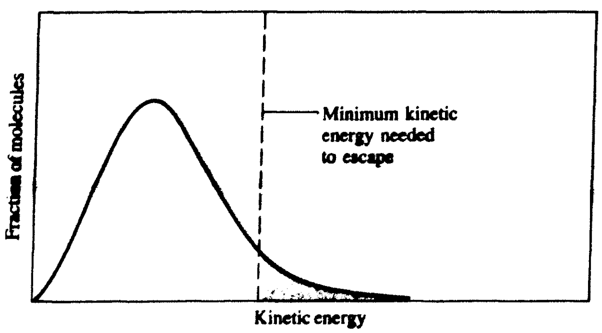


Figure 11.17 Distribution of kinetic energies of surface molecules of a hypothetical liquid compared to the minimum kinetic energy needed to escape from the surface. This minimum energy depends on the magnitude of the attractive forces between molecules. The fraction of molecules having sufficient kinetic energy to escape the liquid is given by the shaded area.

9

Volatility, Vapor Pressure, and Temperature

When vaporization occurs in an open container, as when water evaporates from a bowl, the vapor spreads away from the liquid. Little, if any, is recaptured at the surface of the liquid. Equilibrium never occurs, and the vapor continues to form until the liquid evaporates to dryness. Substances with high vapor pressure (such as gasoline) evaporate more quickly than substances with low vapor pressure (such as motor oil). Liquids that evaporate readily are said to be **volatile**.

Hot water evaporates more quickly than cold water because vapor pressure increases with temperature. As the temperature of a liquid is increased, the molecules move more energetically and can therefore escape more readily from their neighbors. Figure 11.18 depicts the variation in vapor pressure with temperature for four common substances that differ greatly in volatility. Note that in all cases the vapor pressure increases nonlinearly with increasing temperature.

Vapor Pressure and Boiling Point

A liquid boils when its vapor pressure equals the external pressure acting on the surface of the liquid. At this point, bubbles of vapor are able to form within the interior of the liquid. The temperature of boiling increases with increasing external pressure. The boiling point of a liquid at 1 atm pressure is called its **normal boiling point**. From Figure 11.18, we see that the normal boiling point of water is 100°C.

SAMPLE EXERCISE 11.4

Using Figure 11.18, estimate the boiling point of ethanol at 400 mg Hg.

Solution: From Figure 11.18, we see that the boiling point must be about 64°C.

PRACTICE EXERCISE

If we wanted to establish a boiling point of 40°C for diethyl ether, what vapor pressure would we need to maintain in the container? ***Answer:*** about 960 mm Hg

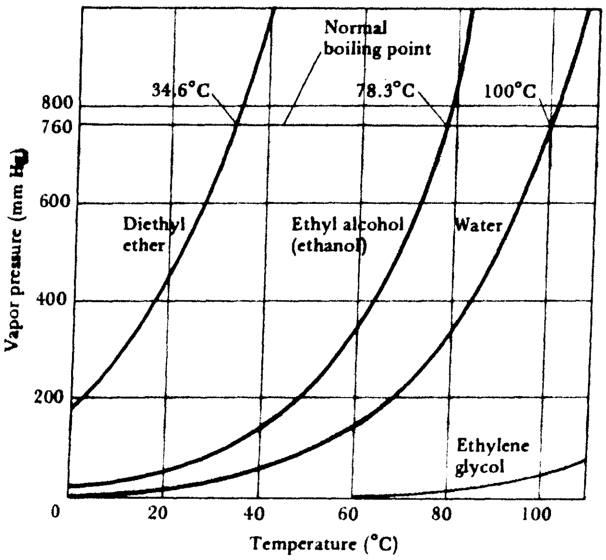


Figure 11.18 Vapor pressure of four common liquids shown as a function of temperature. The temperature at which the vapor pressure is 760 mm Hg is the normal boiling point of each liquid.

CHE30S
Worksheet

NAME ____________________
DATE ____________________

Vapor Pressure and Temperature

1. Using the data in the table below, make a graph of vapor pressure vs temperature on graph paper. Put temperature on the x-axis and vapor pressure on the y-axis. Connect the data for each substance with a smooth curve. Graph all three substances on the same graph. Include a key to identify the three curves.

Temperature (°C) of Three Liquids at Different Vapor Pressures

	10 mm Hg	40 mm Hg	100 mmHg	400 mm Hg	760 mmHg
ethanol	-2.3	19.0	34.9	63.5	78.4
acetone	-31.1	-9.4	7.7	39.5	56.5
cyclohexane	-15.9	6.7	25.5	60.8	80.7

2. Using your graph, determine the following:

   a. the vapor pressure of ethanol at 50 °C.

   b. the vapor pressure of acetone at 50 °C.

   c. the temperature at which the vapor pressure of cyclohexane in 200 mm Hg.

   d. the temperature at which the vapor pressure of acetone in 200 mm Hg.

   e. which of the substances has the highest vapor pressure.

   f. which of the substances has the lowest vapor pressure.

   g. which of the substances would evaporate fastest at room temperature.

   h. which of the substances would evaporate most slowly at room temperature.

3. Which of the substances has the greatest forces of attraction between molecules? Which has the smallest? Explain how you know.

4. Which of the substances would require the most energy to evaporate one mole of the liquid? Which would require the least? Explain how you know.

5. Using your graph, predict the temperature at which the vapor pressure of each substance would be 800 mm Hg.

6. Using your graph, determine the normal boiling temperature of each substance.

7. Using your graph, determine the boiling temperature of each substance at 300 mm Hg.

Date: ____________

Assignment: ____________

From: ____ To: ____ Page No.: ____________

4 Evaporation of water also takes place at temperatures below 100 °C. If a swimmer emerges from water with 90 g of water clinging to him, (a) why does the swimmer feel cold, even though the day may be hot? (b) How much heat energy is necessary to evaporate the water clinging to him?

5 An actively sweating person, such as a laborer or an athlete, on a hot day may lose as much as 3 gallons, or about 11 kg, of water through the skin. (a) How many kilocalories are required to evaporate 11 kg of water? (b) Explain, in terms of regulating body temperature, why sweating is more profuse on hot days than on cold days.

6 Mercury vapor is a substance which is a cumulative poison; that is, repeated small doses of mercury vapor add up to a seriously disabling condition. Explain why, although the boiling point of mercury is 357 °C, it is important to clean up mercury spills thoroughly and promptly.

7 Carbon tetrachloride is another vapor which is poisonous. Its boiling point is 76.8 °C. Explain the following: (a) Carbon tetrachloride poisoning can result from relatively short exposure to its vapor. (b) Spilled carbon tetrachloride can be removed from a room by thorough airing, while spilled mercury must be cleaned up.

8 Fog is composed of tiny droplets of *liquid* water suspended in air. Explain, in terms of the relationship between temperature and the vapor pressure of water, why fog is commonly present during early mornings and late afternoons rather than at the middle of the day.

9 Refer to Figure 5-6 to answer the following questions. (a) What is the normal boiling point of each of the substances listed? (b) At a temperature of 80 °C, which liquids would be boiling? (c) Of all the liquids listed, which has the strongest intermolecular forces?

10 A flask containing ethyl alcohol is attached to a vacuum pump and the pressure is reduced to 12 mm Hg. If the flask is maintained in an ice-water bath at 0 °C, (a) will the alcohol boil? (b) Which, if any, of the other substances listed in Table 5-5 would boil at 0 °C if exposed to this same vacuum pump?

11 Atmospheric pressure falls about 25 mm Hg for every 300 metres above sea level. (a) What would be the boiling temperature of water near the top of Pike's Peak, which has an altitude of about 4,300 metres (over 14,000 feet)? (b) Explain why the comforts of a hot cup of tea or soup are difficult for climbers to obtain near the top of Mount Everest (8,839 metres). (c) Suggest a means of obtaining hot water under such conditions.

12 (a) Explain why food cooks more rapidly in a pressure cooker than in an open pan. (b) Explain why a pressure cooker is more effective in sterilizing baby bottles and surgical instruments than is boiling water in an unpressurized container.

13 What good reason is there for "pressurizing" automobile cooling systems?

14 Liquid sodium has been proposed as a coolant in nuclear power plants to replace water, which is used as a coolant in conventional power plants. Sodium has a melting point of 98 °C and a boiling point of 889 °C. (a) Which would have the higher vapor pressure at 99 °C, sodium or water? (b) Are the intermolecular forces in sodium stronger or weaker than those in water? How can you tell? (c) Which would you expect to have the higher heat of vaporization, liquid sodium or liquid water? Why? (d) Which would you predict to have the higher molar volume in the gaseous phase at, for instance, 1000 °C and 1 atm pressure? Why? (Consider small differences carefully.)

15 Butane, a gas used as a fuel in camping equipment, has a melting point of −138.3 °C and a boiling point of −0.5 °C. Methane, another common fuel, has a melting point of −182 °C and a boiling point of −161 °C. Predict whether the following physical properties will be higher or lower for butane than methane: (a) intermolecular attraction; (b) heat of vaporization; (c) molar heat of melting; (d) vapor pressure at any given temperature.

16 How much heat must be removed from an ice cube tray full of water at 0 °C to freeze it if the tray holds 450 g of water?

17 Explain, in terms of molecular activity, why the molar heat of vaporization of a substance is higher than its molar heat of melting.

18 Consider a stoppered flask partially filled with salt water, in which there is undissolved salt at the bottom of the flask even after several days of shaking and swirling. (a) How many phases are present in the flask? (b) Describe the components of each phase. (c) Which phases are pure substances and which are solutions? (d) How could you separate the solutions into their component pure substances? (e) Would you expect the liquid

will have enough kinetic energy (*i.e.*, are moving fast enough) to overcome the forces of attraction in the liquid and escape into the vapor (Figure 5-5). At higher temperatures there will be more molecules in the vapor phase. Hence, vapor pressure will increase with temperature.

TABLE 5-5 VAPOR PRESSURES OF LIQUIDS

Temp (°C)	Water (mm Hg)	Ethyl Alcohol (mm Hg)	Carbon Tetrachloride (mm Hg)	Methyl Salicylate (mm Hg)	Benzene (mm Hg)
−10	2.1	5.6	19		15
−5	3.2	8.3	25		20
0	4.6	12.2	33		27
5	6.5	17.3	43		35
10	9.2	23.6	56		45
15	12.8	32.2	71		58
20	17.5	43.9	91		74
25	23.8	59.0	114		94
30	31.8	78.8	143		118
35	42.2	103.7	176		147
40	55.3	135.3	216		182
45	71.9	174.0	263		225
50	92.5	222.2	317		271
55	118.0	280.6	379		325
60	149.4	352.7	451	1.41	389
65	190.0	448.8	531	1.90	462
70	233.7	542.5	622	2.52	547
75	289.1	666.1	720	3.40	643
80	355.1	812.6	843	4.41	753
85	433.6	986.7	968	5.90	877
90	525.8	1187	1122	7.63	1020
95	633.9	1420	1270	9.93	1180
100	760.0	1693.3	1463	12.8	1360

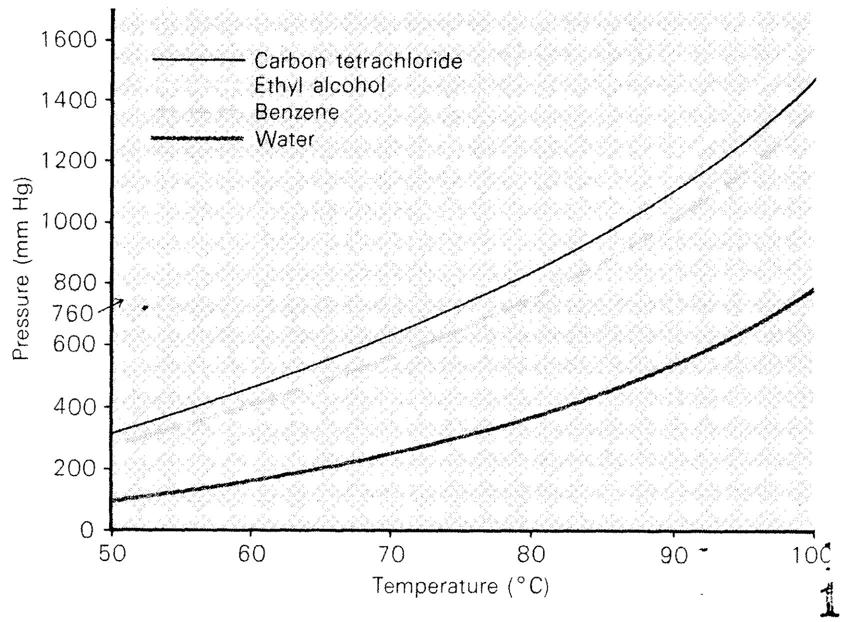
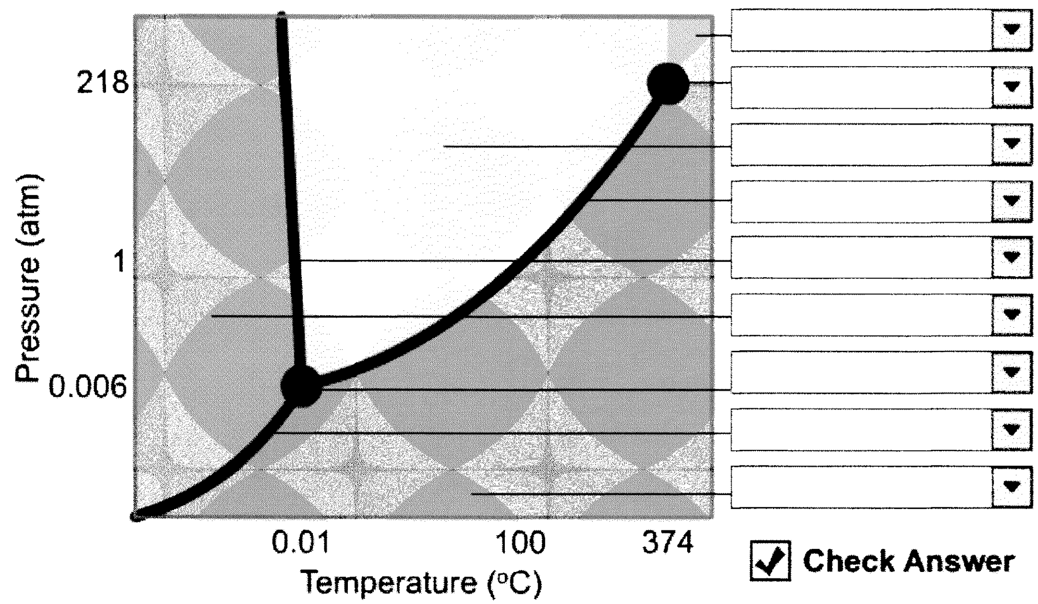


Fig. 5-6 Vapor pressure versus temperature for some common liquids.

Question 1:

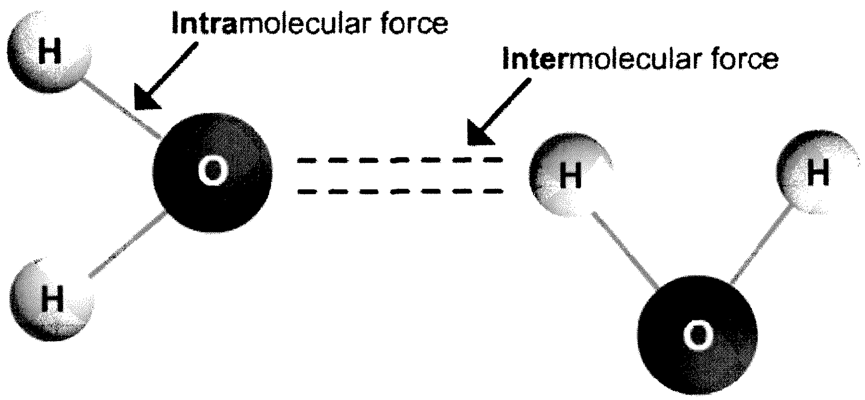
Label the phase diagram for water by selecting from the menus.



INTERMOLECULAR FORCES

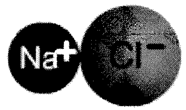
Inter Vs. Intra

Just as an interstate highway runs between states, **intermolecular** forces are attractions between molecules. In contrast, **intramolecular** forces occur within a molecule and include covalent bonds between atoms.

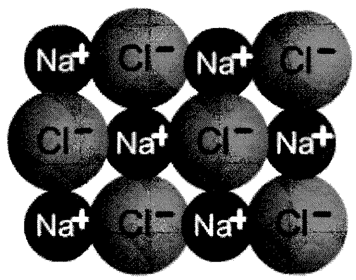


To help you remember, think interstate (or international) versus intrastate commerce.

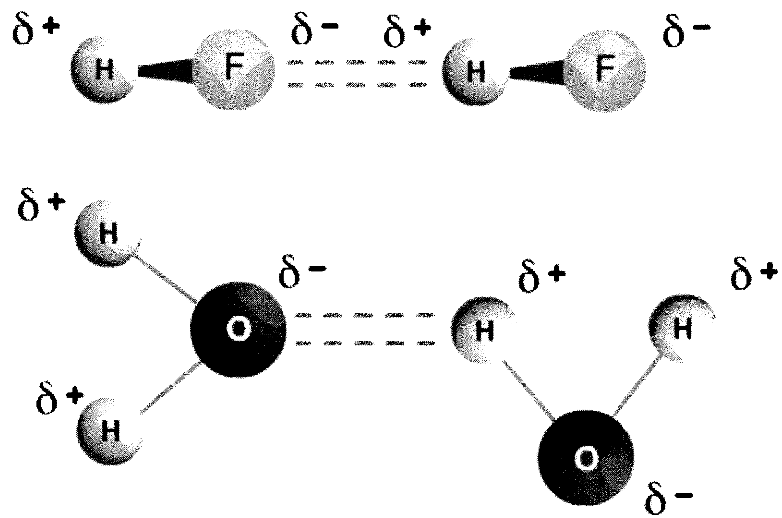
A single ionic molecule, like NaCl, is held together by the attraction between opposite charges. This attraction is called an ionic bond.



A large sample of sodium chloride is also held together by attractions between opposite charges. These attractions, called **ion–ion** forces, are the strongest kind of intermolecular force.



Hydrogen bonding is a special kind of dipole–dipole attraction that occurs when a hydrogen atom is covalently bonded to nitrogen, oxygen, or fluorine.

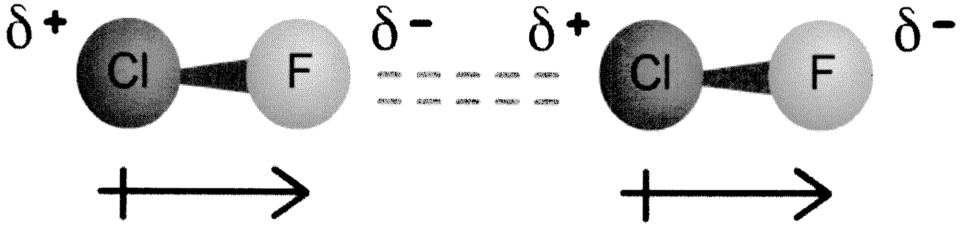


Hydrogen bonding between molecules is the strongest kind of dipole–dipole attraction.

Chapter 10

INTERMOLECULAR FORCES

Dipole-Dipole Forces



As the polarity of the molecules increases, so will the dipole-dipole forces.

Example: FBr > FCl

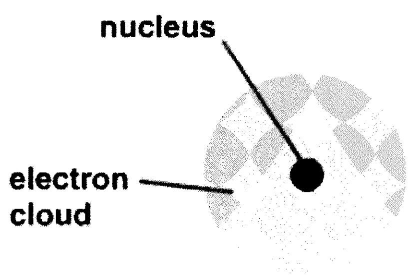
Reason: the electronegativity difference between F and Br is greater than the difference between F and Cl.

Electronegativites:

F	4.0
Cl	3.0
Br	2.8

Dispersion Forces

Dispersion forces, or London forces, occur when the electron clouds of two molecules within close proximity of each other are distorted because of the repulsion between the electrons. The atomic charge distribution is disrupted for a split second, resulting in a brief dipole moment, or **induced dipole**, where one side of the molecule becomes slightly more negative, and the other side relatively positive. This dipole can induce similar dipoles in other nearby molecules. These dipoles fluctuate rapidly.



INTERMOLECULAR FORCES

Relative Strengths

It is important to remember the relative strengths of the different intermolecular forces because these strengths relate directly to melting and boiling points.

Stronger intermolecular forces	=	**Higher melting and boiling points**
Weaker intermolecular forces	=	**Lower melting and boiling points**

Question 1:

Rank the following substances from strongest to weakest intermolecular forces by dragging each substance to the correct box.

He NH_3 NF_3 NaCl

[ ] > [ ] > [ ] > [ ] **Check Answer**

Question 2:

Rank the following substances from strongest to weakest intermolecular forces by dragging each substance to the correct box.

HF F_2 FCl

[ ] > [ ] > [ ]

Check Answer

Question 3:

Rank the following substances from strongest to weakest intermolecular forces by dragging each substance to the correct box.

$NaCl$ $MgCl_2$ $AlCl_3$ MgS $NaBr$

☐ > ☐ > ☐ > ☐ > ☐ ☑ **Check Answer**

Intermolecular Forces

For questions 1-5, identify the main type of intermolecular force in each compound:

1) carbon disulfide

2) ammonia

3) oxygen

4) CH_2F_2

5) C_2H_6

Rank the following compounds by increasing melting point:

6) C_2H_6, C_2H_5OH, C_2H_5F

7) H_2S, H_2O, H_2

8) BBr_3, BI_3, BCl_3

Intermolecular Forces Worksheet

1) Using your knowledge of molecular structure, identify the main intermolecular force in the following compounds. You may find it useful to draw Lewis structures to find your answer.

   a) PF_3 ______________________________

   b) H_2CO ______________________________

   c) HF ______________________________

2) Explain how dipole-dipole forces cause molecules to be attracted to one another.

3) Rank the following compounds from lowest to highest boiling point: calcium carbonate, methane, methanol (CH_4O), dimethyl ether (CH_3OCH_3).

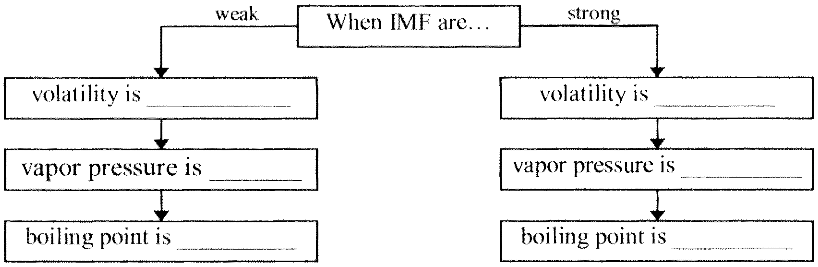
4) Explain why nonpolar molecules usually have much lower surface tension than polar ones.

Name ______________________________ Period ________ Date ___________

Changes of State

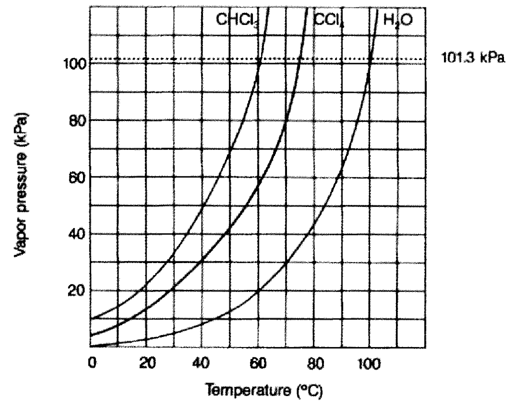
PART A – INTERMOLECULAR FORCES

1. Fill in the diagram (with high or low) to show how intermolecular forces influence the **volatility**, **vapor pressure**, and **boiling point** of a substance.



PART B – VAPOR PRESSURE GRAPHS

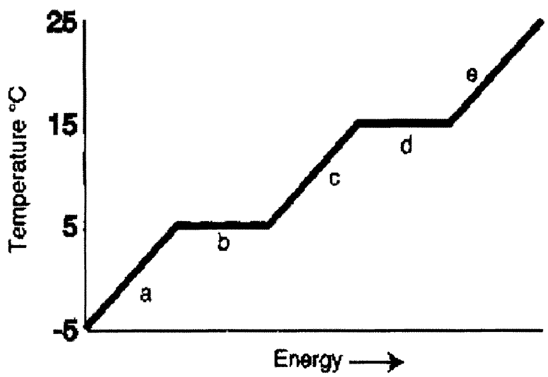
Use the graph below to answer the following questions.



2. What is the vapor pressure of $CHCl_3$ at 50°C? ____________
3. What is the boiling point of H_2O when the external pressure is 30 kPa? ____________
4. What is the normal boiling point of CCl_4? ____________
5. Which substance has the weakest IMF? ____________

PART C – HEATING CURVES.

Use the heating curve below to answer the following questions.



6. What is the melting point of the substance? ______
7. What is the boiling point of the substance? ______
8. Which letter represents heating of the solid? ______
9. Which letter represents heating of the vapor? ______
10. Which letter represents melting of the solid? ______
11. Which letter represents boiling of the liquid? ______

PART D – PHASE DIAGRAMS.

Use the phase diagram for water below to answer the following questions.

12. What is the state of water at 2 atm and 50°

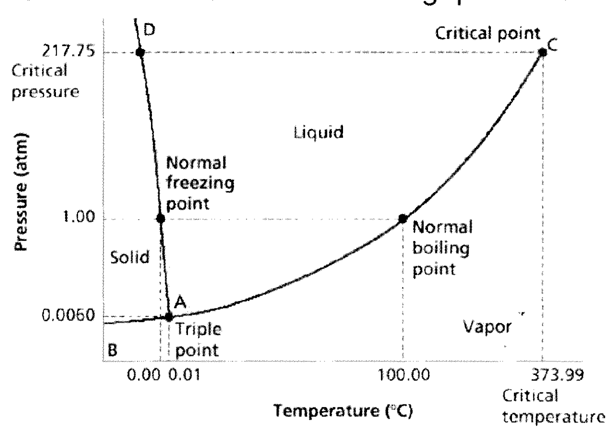
    ____

13. What phase change will occur if the temperature is lowered from 80°C to -5°C at 1 atm?

    ____

14. You have ice at -10°C and 1 atm. What could you do in order cause the ice to sublime?

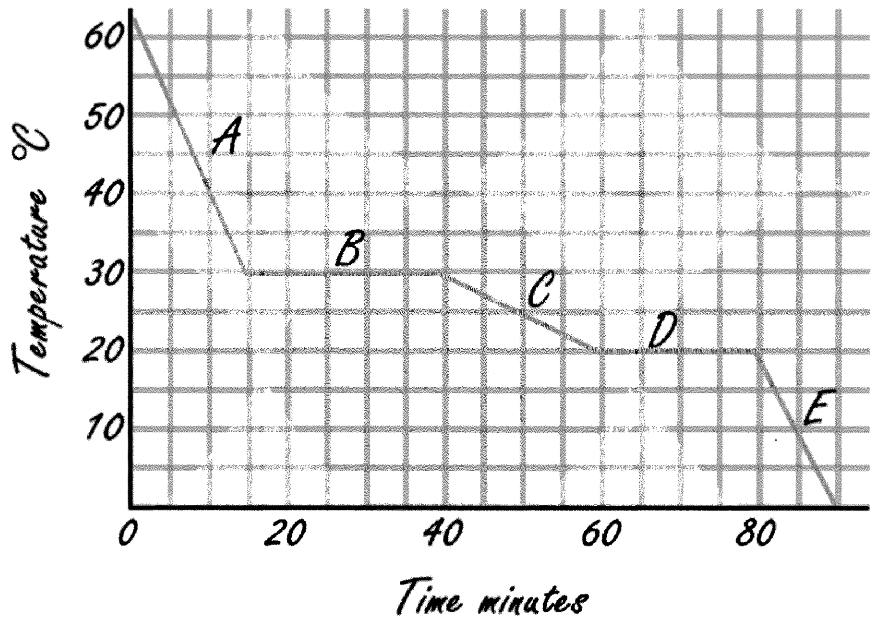
    ____



29

Chemistry 30S: Physical Properties of Matter Review Questions

1. Answer the following questions based on the graph of temperature versus time for mysterious substance X.



a. Would this curve be classified as a warming curve or a cooling curve?

b. Does the graph represent an endothermic or an exothermic process? Explain.

c. What is the melting point and boiling point of substance X?

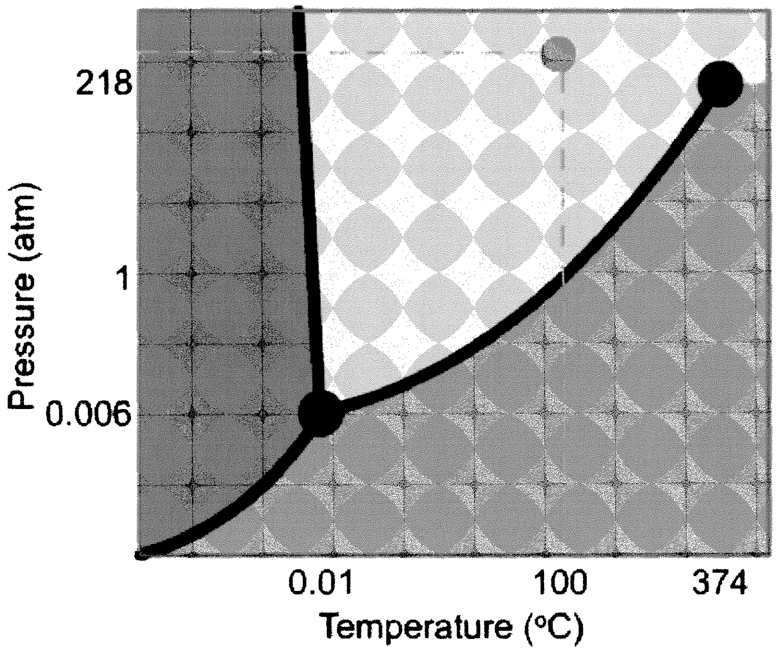
d. At which point(s) on the graph is the kinetic energy constant?

e. At each plateau the substance continues to lose heat even though there is no change in temperature. Explain what is happening at the particle level.

Chemistry 30S: Physical Properties of Matter Review Questions

2. Place the following labels in the appropriate areas on the phase diagram.

A. Solid	B. Triple point	C. Supercritical fluid
D. Liquid	E. Critical temperature	F. Solid & gas in equilibrium
G. Gas	H. Critical pressure	I. Vaporization



3. Answer the following questions based on the phase diagram.
   a. This phase diagram likely represents which substance? Explain your reasoning.

____

____

____

   b. What is the boiling point of the substance at 0.006 atm?

   c. Will the substance undergo sublimation at standard atmospheric pressure? Explain your reasoning.

____

____

____

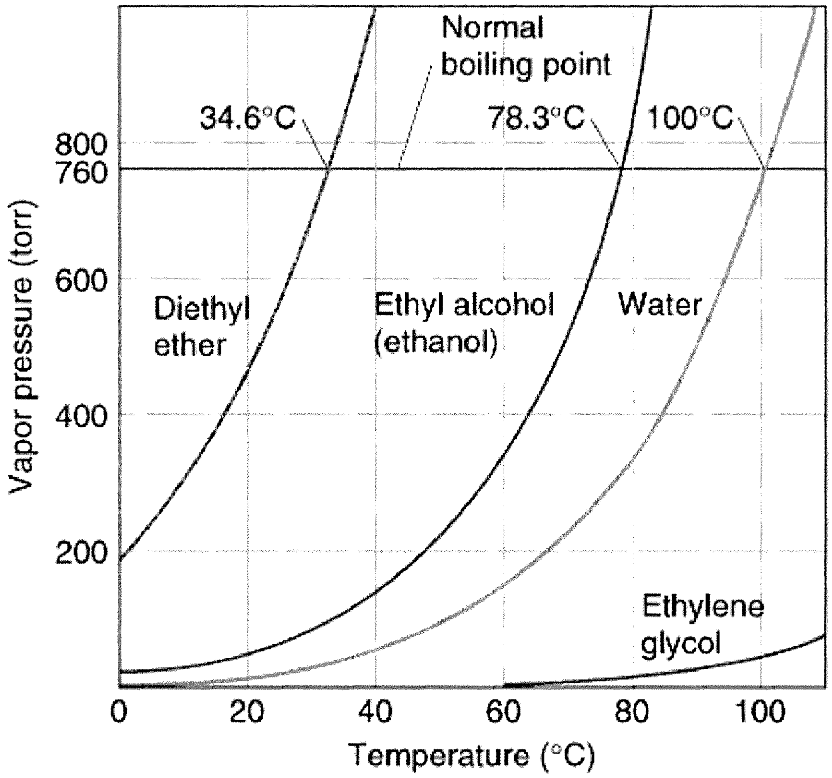
Chemistry 30S: Physical Properties of Matter Review Questions

4. Use the curve of pressure versus temperature to answer the following questions.
   a. Which substance has the greatest intermolecular forces?

   b. Which substance requires the least amount of energy to vaporize?

   c. At 600 torr (1 torr = 1 mm Hg) and 80°C which substance(s) is/are in the gas phase?

   d. At what pressure would water boil at 78.3°C? Would ethanol be a liquid or a gas under these same conditions?



5. Atmospheric pressure drops by 25 mm Hg for every 300 m in elevation.
   a. Calculate the temperature at which water boils at the peak of a mountain with an elevation of 6,720 m.

b. Calculate the boiling point of water at 480 m below sea level.

6. Explain the difference between evaporation and boiling.

7. State whether the following processes are exothermic or endothermic.

   a. A campfire burning
   b. Snow melting
   c. $C_3H_{8(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$ + heat
   d. $H_{2(g)} + O_{2(g)} \rightarrow H_2O_{(l)}$ $\Delta H = +150$ kJ/mol

8. Which is more severe; a steam burn or a scald from hot water? Explain your reasoning.

9. Use the principles of vapor pressure and heat transfer to explain why a pressurized car radiator uses a liquid that is tightly capped in container with a very high surface area to prevent the engine from overheating.