Solutions and Intermolecular Forces – HW

PSI Chemistry

Name_________________________________

Solutions-HW

Section A&B

1. In a saturated solution of salt water, ____________.
   A) The rate of crystallization > the rate of solution.
   B) The rate of solution > the rate of crystallization
   C) Seed crystal addition may cause massive crystallization
   D) The rate of crystallization = the rate of solution
   E) Addition of more water causes massive crystallization

2. Of the following, ____________ should be immiscible with carbon tetrachloride, CCl₄.
   A) C₆H₁₄
   B) Br₂
   C) CH₃CH₂OH
   D) C₃H₈
   E) I₂

3. Which one of the following substances is more likely to dissolve in benzene (C₆H₆)?
   A) CH₃CH₂OH
   B) NH₃
   C) NaCl
   D) CCl₄
   E) HBr

4. Which one of the following is most soluble in water?
   A) CH₃OH
   B) CH₃CH₂CH₂OH
   C) CH₃CH₂OH
   D) CH₃CH₂CH₂CH₂OH
   E) CH₃CH₂CH₂CH₂CH₂OH

5. Which one of the following is most soluble in hexane (C₆H₁₄)?
   A) CH₃OH
   B) CH₃CH₂CH₂OH
   C) CH₃CH₂OH
   D) CH₃CH₂CH₂CH₂OH
   E) CH₃CH₂CH₂CH₂CH₂OH

6. An unsaturated solution is one that ____________.
   A) Has no double bonds
   B) Contains the maximum concentration of solute possible and is in equilibrium with undissolved solute.
   C) Has a concentration lower than the solubility
   D) Contains more dissolved solute than the solubility
   E) Contains no solute

7. A solution with a concentration higher than the solubility ____________.
   A) Is not possible
B) Is unsaturated
C) Is superficial
D) Is saturated
E) Is supersaturated

8. The phrase "like dissolves like" refers to the fact that __________.
   A) gases can only dissolve other gases
   B) polar solvents dissolve polar solutes; nonpolar solvents dissolve nonpolar solutes
   C) solvents can only dissolve solutes of similar molar mass
   D) condensed phases can only dissolve other condensed phases
   E) polar solvents dissolved nonpolar solutes and vice versa

9. The process of solute particles being surrounded by solvent particles is known as _____.
   A) salutation
   B) agglomeration
   C) solvation
   D) agglutination
   E) dehydration

10. A saturated solution __________.
    A) contains as much solvent as it can hold
    B) contains no double bonds
    C) contains dissolved solute in equilibrium with undissolved solute
    D) will rapidly precipitate if a seed crystal is added
    E) cannot be attained

11. An unsaturated solution __________.
    A) has no double bonds
    B) contains the maximum concentration of solute possible, and is in equilibrium with undissolved solute
    C) has a concentration lower than the solubility
    D) contains more dissolved solute than the solubility allows
    E) contains no solute

12. A solution with a concentration higher than the solubility is __________.
    A) is not possible
    B) is unsaturated
    C) is supercritical
    D) is saturated
    E) is supersaturated

13. A supersaturated solution __________.
    A) is one with more than one solute
    B) is one that has been heated
    C) is one with a higher concentration than the solubility
    D) must be in contact with undissolved solid
    E) exists only in theory and cannot actually be prepared
14. A sample of potassium nitrate (49.0 g) is dissolved in 101 g of water at 100 °C, with precautions taken to avoid evaporation of any water. The solution is cooled to 30.0 °C and no precipitate is observed. This solution is __________.
   A) hydrated     B) placated     C) saturated     D) unsaturated     E) supersaturated

15. A sample of potassium chlorate (15.0 g) is dissolved in 201 g of water at 70 °C with precautions taken to avoid evaporation of any water. The solution is cooled to 30.0 °C and no precipitate is observed. This solution is __________.
   A) hydrated     B) miscible     C) saturated     D) unsaturated     E) supersaturated

16. A sample of potassium nitrate (49.0 g) is dissolved in 101 g of water at 100 °C with precautions taken to avoid evaporation of any water. The solution is cooled to 30.0 °C and a small amount of precipitate is observed. This solution is __________.
   A) hydrated     B) placated     C) saturated     D) unsaturated     E) supersaturated

Section C

17. Which one of the following substances would be the most soluble in CCl₄?
   A) CH₃CH₂OH     B) H₂O     C) NH₃     D) C₁₀H₂₂     E) NaCl

18. Which of the following substances is more likely to dissolve in water?
   A) HOCH₂CH₂OH
   B) CHCl₃
   C) CH₃(CH₂)₉HC=O
   D) CH₃(CH₂)₇CH₂OH
   E) CCl₄

19. Which of the following substances is more likely to dissolve in CH₃OH ?
   A) CCl₄     B) Kr     C) N₂     D) CH₃CH₂OH     E) H₂

20. Which one of the following substances is more likely to dissolve in CCl₄?
   A) CBr₄     B) HBr     C) HCl     D) CH₃CH₂OH     E) NaCl

21. Which one of the following substances is more likely to dissolve in benzene (C₆H₆)?
   A) CH₃CH₂OH     B) NH₃     C) NaCl     D) CCl₄     E) HBr
22. Which one of the following is most soluble in water?
   A) CH₃OH
   B) CH₃CH₂CH₂OH
   C) CH₃CH₂OH
   D) CH₃CH₂CH₂CH₂OH
   E) CH₃CH₂CH₂CH₂CH₂OH

23. Which one of the following is most soluble in hexane (C₆H₁₄)?
   A) CH₃OH
   B) CH₃CH₂CH₂OH
   C) CH₃CH₂OH
   D) CH₃CH₂CH₂CH₂OH
   E) CH₃CH₂CH₂CH₂CH₂OH

24. Pressure has an appreciable effect on the solubility of __________ in liquids.
   A) gases      B) solids      C) liquids      D) salts      E) solids and liquids

Section D
25. An aqueous solution contains 28% phosphoric acid by mass. This means that
   __________.
   A) 1 mL of this solution contains 28 g of phosphoric acid
   B) 1 L of this solution has a mass of 28 g
   C) 100 g of this solution contains 28 g of phosphoric acid
   D) 1 L of this solution contains 28 mL of phosphoric acid
   E) the density of this solution is 2.8 g/mL

26. A solution is prepared by dissolving 25 g of CaCl₂ in 475 g of water. The density of the
    resulting solution is 1.05 g/mL. The concentration of CaCl₂ is __________% by mass.
    A) 5.00    B) 1.25    C) .01    D) 0.05    E) 6.25

27. The concentration of urea in a solution prepared by dissolving 16 g of urea in 64 g of
    H₂O is __________% by mass.
    A) 25    B) 41    C) 0.25    D) 0.41    E) 0.48

28. A solution contains 11% by mass of sodium chloride. This means that __________.
    A) there are 11 g of sodium chloride in in 1.0 mL of this solution
    B) 100 g of the solution contains 11 g of sodium chloride
    C) 100 mL of the solution contains 11 g of sodium chloride
    D) the density of the solution is 11 g/mL
    E) the molality of the solution is 11

29. A solution is prepared by dissolving 17.0 g of NH₃ in 360.0 g of water. The mole
    fraction of NH₃ in the solution is __________.
    A) 0.064    B) 0.05    C) 0.940    D) 0.92    E) 16.8

30. The mole fraction of He in a gaseous solution prepared from 4.0 g of He, 20 g of Ar, and
    10.0 g of Ne is __________.
    A) 0.50    B) 1.5    C) 0.20    D) 0.11    E) 0.86

31. The mole fraction of urea (MW = 60.0 g/mol) in a solution prepared by dissolving 15 g of
    urea in 40.5 g of H₂O is __________.
32. Calculate the mole fraction of HCl in a 10.0\% (by mass) aqueous solution.
   A) 0.00111  B) 0.0344  C) 0.0520  D) 0.0548  E) 0.122

33. A solution contains equal masses of glucose (molecular mass 180) and toluene (molecular mass 90). What is the mole fraction of glucose in the solution?
   A) 1/4  B) 1/3  C) 1/2  D) 2/3  E) 3/4

Section E
34. Which one of the following concentration units varies with temperature?
   A) molarity  B) mass percent  C) mole fraction  D) molality  E) all of the above

35. Which one of the following is a correct expression for molarity?
   A) mol solute/L solvent
   B) mol solute/mL solvent
   C) mmol solute/mL solution
   D) mol solute/kg solvent
   E) μmol solute/L solution

36. Which one of the following is not true concerning 2.00 L of 0.100 M solution of \( \text{Ca}_3(\text{PO}_4)_2 \)?
   A) This solution contains 0.200 mol of \( \text{Ca}_3(\text{PO}_4)_2 \).
   B) This solution contains 0.800 mol of oxygen atoms.
   C) 1.00 L of this solution is required to furnish 0.300 mol of \( \text{Ca}^{2+} \) ions.
   D) There are \( 6.02 \times 10^{23} \) phosphorus atoms in 500.0 mL of this solution.
   E) This solution contains \( 6.67 \times 10^{-2} \) mol of \( \text{Ca}^{2+} \).

37. When 0.500 mol of HC\(_2\)H\(_3\)O\(_2\) is combined with enough water to make a 500.0 mL solution, the concentration of HC\(_2\)H\(_3\)O\(_2\) is _________ M.
   A) 3.33  B) 1.00  C) 0.835  D) 0.00167  E) 0.150

38. What is the concentration (M) of CH\(_3\)OH in a solution prepared by dissolving 16 g of CH\(_3\)OH in sufficient water to give exactly 250 mL of solution?
   A) 11.9  B) 1.59 \times 10^{-3}  C) 0.0841  D) 2.00  E) 11.9 \times 10^{-3}

39. What is the concentration (M) of a NaCl solution prepared by dissolving 9.3 g of NaCl in sufficient water to give 350 mL of solution?
   A) 18  B) 0.16  C) 0.45  D) 27  E) 2.7 \times 10^{-2}

40. How many grams of CH\(_3\)OH must be added to water to prepare 150 mL of a solution that is 2.0 M CH\(_3\)OH?
   A) 9.6 \times 10^{3}  B) 4.3 \times 10^{2}  C) 2.4  D) 9.6  E) 4.3

41. The molarity (M) of an aqueous solution containing 22.5 g of sucrose (C\(_{12}\)H\(_{22}\)O\(_{11}\)) in 35.5 mL of solution is __________.
   A) 0.0657  B) 1.85 \times 10^{-3}  C) 1.85  D) 3.52  E) 0.104

42. The molarity (M) of an aqueous solution containing 52.5 g of sucrose (C\(_{12}\)H\(_{22}\)O\(_{11}\)) in 35.5 mL of solution is __________.
   A) 5.46  B) 1.48  C) 0.104  D) 4.32  E) 1.85
43. How many grams of NaOH (MW = 40.0) are there in 500.0 mL of a 0.175 M NaOH solution?
   A) 2.19 x 10^{-3}  B) 114  C) 14.0  D) 3.50  E) 3.50 x 10^{3}

44. How many grams of H₃PO₄ are in 175 mL of a 3.5 M solution of H₃PO₄?
   A) 0.61  B) 60  C) 20  D) 4.9  E) 612

45. How many grams of sodium chloride are there in 55.0 mL of a 1.90 M aqueous solution of sodium chloride?
   A) 0.105  B) 6.11  C) 3.21  D) 6.11 x 10^{3}  E) 12.2

46. How many grams of sodium chloride are there in 550.0 mL of a 1.90 M aqueous solution of sodium chloride?
   A) 61.1  B) 1.05  C) 30.5  D) 9.6 x 10^{4}  E) 122

47. A substance is dissolved in water, forming a 0.50 molar solution. If 4.0 liters of solution contains 240 grams of the substance, what is the molecular mass of the substance?
   A) 60 g/mol  B) 120 g/mol  C) 240 g/mol  D) 480 g/mol  E) 640 g/mol

48. The concentration of nitrate ion in a solution that contains 0.900 M aluminum nitrate is __________.
   A) 0.900  B) 0.450  C) 0.300  D) 2.7  E) 1.80

49. An aqueous solution of a soluble compound (a nonelectrolyte) is made by dissolving 33.2 g of the compound in a water to form 250 mL solution. The solution has an osmotic pressure of 1.2 atm at 25°C. What is the molar mass of the compound?
   A) 1.0 x 10^{3}  B) 2.7 x 10^{3}  C) 2.3 x 10^{2}  D) 6.8 x 10^{2}  E) 28

51. How many moles of Na₂SO₄ must be added to 500 milliliters of water to create a solution that has a 2 molar concentration of the Na⁺ ion? (assume the volume of the solution does not change.)
   A) 0.5  B) 1  C) 2  D) 4  E) 5

52. If 14 g of MgBr₂ (molar mass 184 grams) is dissolved in water to form 0.50 liters of solution, what is the concentration of bromine ion in the solution?
   A) 1.0 M  B) 0.50 M  C) 0.25 M  D) 2.0 M  E) 4.0 M

Section F

53. Molality is defined as the __________.
   A) moles solute/moles solvent  B) moles solute/Liters solution  C) moles solute/kg solution  D) moles solute/kg solvent  E) none (dimensionless)

54. Of the concentration units below, only __________ is temperature dependent.
   A) mass %  B) ppm  C) ppb  D) molarity  E) molality
55. The concentration of KBr in a solution prepared by dissolving 2.21 g of KBr in 897 g of water is ________ molal.
   A) 2.46 B) 0.0167 C) 0.0207 D) 2.07 x 10^{-5} E) 0.0186

56. The concentration of lead nitrate (Pb(NO$_3$)$_2$) in a 0.726 M solution is ________ molal. The density of the solution is 1.202 g/mL.
   A) 0.476 B) 1.928 C) 0.755 D) 0.819 E) 0.650

57. The concentration of a benzene solution prepared by mixing 12.0 g C$_6$H$_6$ with 38.0 g CCl$_4$ is ________ molal.
   A) 4.04 B) 0.240 C) 0.622 D) 0.316 E) 0.508

58. A solution is prepared by dissolving 23.7 g of CaCl$_2$ in 375 g of water. The density of the resulting solution is 1.05 g/mL. The concentration of CaCl$_2$ in this solution is ________ molal.
   A) 0.214 B) 0.569 C) 5.70 D) 63.2 E) 1.76

59. The concentration of HCl in a solution that is prepared by dissolving 5.5 g of HCl in 1100 g of C$_2$H$_6$O is ________ molal.
   A) 27.5 B) 7.5 x 10^{-4} C) 3.3 x 10^{-2} D) 5 E) 1.3

60. The concentration of urea (MW = 60.0 g/mol) in a solution prepared by dissolving 30 g of urea in 50 g of H$_2$O is ________ molal.
   A) 96 B) 10 C) 0.68 D) 6.3 E) 0.11

61. The concentration of KBr in a solution prepared by dissolving 6 g of KBr in 897 g of water is ________ molal.
   A) 2.46 B) 0.0167 C) 0.0207 D) 2.07 x 10^{-5} E) 0.0186

Section G

62. As the concentration of a solute in a solution increases, the freezing point of the solution ________ and the vapor pressure of the solution ________.
   A) increases, increases B) increases, decreases C) decreases, increases D) decreases, decreases E) decreases, is unaffected

63. Which of the following liquids will have the lowest freezing point?
   A) pure H$_2$O
   B) 0.10 m aqueous glucose
   C) 0.15 m aqueous glucose
   D) 0.20 m aqueous glucose
   E) 0.25 m aqueous glucose

64. Which of the following aqueous solutions will have the highest boiling point?
   A) 0.10 m NaCl
   B) 0.15 m NaCl
   C) 0.20 m NaCl
   D) 0.25 m NaCl
   E) pure water

65. Which produces the greatest number of ions when one mole dissolves in water?
   A) NaCl B) NH$_4$NO$_3$ C) NH$_4$Cl D) Na$_2$SO$_4$ E) C$_{12}$H$_{22}$O$_{11}$
66. Which of the following liquids will have the lowest freezing point?
   A) pure H₂O
   B) 0.10 m aqueous KBr
   C) 0.10 m aqueous FeBr₂
   D) 0.10 m aqueous FeBr₃
   E) 0.10 m aqueous CaBr₂

67. Of the following, a 0.1 M aqueous solution of _________ will have the lowest freezing point.
   A) NaCl   B) Al(NO₃)₃   C) K₂CrO₄   D) Na₂SO₄   E) sucrose, C₁₂H₂₂O₁₁

68. Of the following, a 0.2 M aqueous solution of _________ will have the highest freezing point.
   A) (NH₄)₃PO₄   B) Pb(NO₃)₂   C) Na₃PO₄   D) Mg(NO₃)₂   E) NaCl

69. Which of the following aqueous solutions will have the highest boiling point?
   A) 0.10 m Na₂SO₄
   B) 0.10 m glucose, C₆H₁₂O₆
   C) 0.10 m sucrose, C₁₂H₂₂O₁₁
   D) 0.10 m NaCl
   E) 0.10 m CuSO₄

70. Colligative properties of solutions include all of the following except _________.
   A) depression of vapor pressure upon addition of a solute to a solvent
   B) elevation of the boiling point of a solution upon addition of a solute to a solvent
   C) depression of the freezing point of a solution upon addition of a solute to a solvent
   D) an increase in the osmotic pressure of a solution upon the addition of more solute
   E) the increase of reaction rates with increase in temperature

71. The vapor pressure of pure water at 25°C is 23.8 torr. What is the vapor pressure (torr) of water above a solution prepared by dissolving 18 g glucose (a nonelectrolyte, MW = 180 g/mol) in 95 g of water?
   A) 24.3    B) 23.4    C) 0.451    D) 0.443    E) 23.8

72. Which liquid will have the lowest freezing point?
   A) pure H₂O
   B) aq. 0.60 m glucose
   C) aq. 0.60 m sucrose
   D) aq. 0.24 m FeI₃
   E) aq. 0.50 m KF

73. The freezing point of ethanol (C₂H₅OH) is -114.6°C. The molal freezing point depression constant for ethanol is 2.00°C/m. What is the freezing point (°C) of a solution prepared by dissolving 50.0 g of glycerin (C₃H₈O₃, a nonelectrolyte) in 200 g of ethanol?
   A) -115    B) -5.42    C) -132.3    D) -120.0    E) -114.6

74. Which produces the greatest number of ions when one mole dissolves in water?
75. For an aqueous solution of a nonvolatile compound, the vapor pressure will be ________, the boiling point will be ________, and the freezing point will be ________ than pure water.
A) lower, lower, lower
B) lower, higher, lower
C) lower, higher, higher
D) higher, higher, lower
E) higher, lower, higher

76. Which of the aqueous solutions will have the highest boiling point?
A) 0.10 m Na₂SO₄
B) 0.20 m glucose
C) 0.25 m sucrose
D) 0.10 m NaCl
E) 0.10 m SrSO₄

77. Calculate the freezing point of a 0.05500 m aqueous solution of NaNO₃.
A) 0.0286˚C
B) -0.1023˚C
C) 0.1023˚C
D) -0.05627˚C
E) -0.2046˚C

78. A solution prepared by dissolving 0.60 g of nicotine (a nonelectrolyte) in water to make 12 mL of solution has an osmotic pressure of 7.55 atm at 25˚C. The molecular weight of nicotine is ________ g/mol.
A) 28
B) 43
C) 50
D) 160
E) 0.60

79. A solution prepared by dissolving 6.00 g of an unknown nonelectrolyte in enough water to make 1.00 L of solution. The osmotic pressure of this solution is 0.750 atm at 25.0˚C. What is the molecular weight (g/mol) of the unknown solute?
A) 16.4
B) 196
C) 110
D) 30.6
E) 5.12 x 10⁻³

80. Which of the following is (are) Colligative properties?
   I. Freezing point depression
   II. Vapor pressure lowering
   III. Boiling point elevation
A) I only
81. When 31 grams of a nonionic substance is dissolved in 2.00 kg water, the observed freezing point depression of the solution is 0.93˚C. If k_f for water is 1.86˚C/m, which of the following expressions is equal to the molar mass of the substance?

A) \((31.0)(0.93)(2.00)\) g/mol \((1.86)\)

B) \((31.0)(1.86)\) g/mol \((0.93)(2.00)\)

C) \((2.00)(1.86)\) g/mol \((31.0)(0.93)\)

D) \((0.93)\) g/mol \((31.0)(0.93)(2.00)\)

E) \((31.0)(0.93)(1.86)(2.00)\) g/mol

82. What is the boiling point of a 2 M solution of NaCl in water? (The boiling point elevation constant, k_b, for water is 0.5˚C/m)

A) 100˚C

B) 101˚C

C) 102˚C

D) 103˚C

E) 104˚C

83. When an aqueous salt solution is compared to water, the salt solution will have

A) a higher boiling point, a lower freezing point, and a lower vapor pressure.

B) a higher boiling point, a higher freezing point, and a lower vapor pressure.

C) a higher boiling point, a higher freezing point, and a higher vapor pressure.

D) a lower boiling point, a lower freezing point, and a lower vapor pressure.

E) a lower boiling point, a higher freezing point, and a higher vapor pressure.

**ANSWERS**

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Solutions Free Response Practice

1. An aqueous solution with a volume of 100 mL, containing 0.122 g of an unknown molecular compound, has an osmotic pressure of 16.0 torr at 20.0°C. What is the molar mass of the solute?

2. A 0.562 g sample of an unknown substance was dissolved in 17.4 g benzene. The freezing point of the solution was 4.075°C. The freezing point of pure benzene is 5.455°C. For benzene, \( K_f = 5.065°C/m \), and \( K_b = 2.6°C/m \). Assume that the solute is a nonelectrolyte.
   a. What is the molality of the solution?
   b. What is the molar mass of the unknown?
   c. If the boiling temperature of pure benzene is 80.2°C, what is the boiling temperature of the solution?
   d. What is the van't Hoff factor? Explain whether the use of one was necessary in these calculations.

3. Ethylene glycol (\( C_2H_4(OH)_2 \); 150 grams) is added to ethanol (\( C_2H_5OH \); 250 grams).
   a. Calculate the mass % of ethylene glycol in the solution.
   b. Calculate the molality of ethylene glycol in the solution.
   c. Calculate the mole fraction of ethylene glycol in the solution.

4. In order to depress the freezing point of water to -12°C, how much magnesium nitrate would you have to add to 500 grams of water? Assume that the van't Hoff factor \( i \) is the ideal value. \( K_f \) for water is -1.86 K/m.
Solutions—Free Response Answers

1. \(0.122 \text{ g} = w\)
   
   \(P = 16 \text{ torr} = 16/760 \text{ atm} = 0.021 \text{ atm}\)
   
   \(T = 20 + 273 = 293 \text{ K}\)
   
   \(M = ?\)
   
   \(PV = nRT\)
   
   \(0.021 \times .100 = (0.122/M)(0.0821)(293)\)
   
   \(M = 1397.49\)

2. \(0.560 \text{ g unknown}\)
   
   \(T_{\text{solution}} = 4.075^\circ \text{C}\)
   
   \(17.4 \text{ g Benzene}\)
   
   \(T_{\text{Benzene}} = 5.455\)
   
   \(\Delta T = 5.455 - 4.075 = 1.380\)
   
   \(k_f = 5.065\)
   
   \(k_b = 2.6\)
   
   a. \(\Delta T = \text{molality} \times k_f \rightarrow 1.380 = m \times 5.065 \rightarrow m = 0.2725\)
   
   b. \(M = (1000)(k_f)(w) / (\Delta T)(W) = (1000 \times 5.065 \times 0.562) / (1.38 \times 17.4) = 118.5\)
   
   c. \(\Delta T = (1000)(k_f)(w) / (M)(W) = (1000 \times 2.6 \times 0.562) / (118.5 \times 17.4) = 0.784\)
   
   d. \(i = 1\) Since it is a nonelectrolyte it is considered to be a nonvolatile solute and will not dissociate.

3. Glycol: \(150 \text{ g} \text{ added to} 250 \text{ g}\)
   
   a. \(\text{Mass} \% = 150/400 \times 100 = 37.5\%\)
   
   b. \(\text{Molality} = (\text{mol of glycol}) / (\text{kg of solvent})\)
      
      Glycol: \(62 \text{ g/mol} \rightarrow 2.419 \text{ mol}\)
      
      Ethanol: \(250 \text{ g} = 0.25 \text{ kg}\)
      
      \(m = 2.419/0.025 = 96.76\)
      
   c. \(\text{Mole fraction of glycol} = \text{mol}_{\text{glycol}} / \text{mol}_{\text{total}}\)
      
      \(\text{MF}_{\text{glycol}} = 2.419 / 7.854 = 0.30799\)

4. \(\Delta T = 12 = 1000(k_f)(w) / \text{MW}\)
   
   \(12 = (1000 \times 1.86 \times w) / (86 \times 500)\)
   
   \(w = 277.419\)
I. Intermolecular (IM) Forces

1) The strongest interparticle attractions exist between particles of a __________ and the weakest interparticle attractions exist between particles of a __________.
A) solid, liquid
B) solid, gas
C) liquid, gas
D) liquid, solid
E) gas, solid

2) Which one of the following exhibits dipole-dipole attraction between molecules?
A) XeF₄
B) AsH₃
C) CO₂
D) BCl₃
E) Cl₂

3) Hydrogen bonding is a special case of __________.
A) London-dispersion forces
B) ion-dipole attraction
C) dipole-dipole attractions
D) ion-ion interactions
E) none of the above

4) When NaCl dissolves in water, aqueous Na⁺ and Cl⁻ ions result. The force of attraction that exists between Na⁺ and H₂O is called a(n) __________ interaction.
A) dipole-dipole
B) ion-ion
C) hydrogen bonding
D) ion-dipole
E) London dispersion force

5) Of the following substances, only __________ has London dispersion forces as its only intermolecular force.
A) CH₃OH
B) NH₃
C) H₂S
D) CH₄
E) HCl

6) What is the predominant intermolecular force in CBr₄?
A) London-dispersion forces
B) ion-dipole attraction
C) ionic bonding
D) dipole-dipole attraction
E) hydrogen-bonding

7) Of the following substances, only __________ has London dispersion forces as the only intermolecular force.
A) CH₃OH
B) NH₃
C) H₂S
D) Kr
E) HCl

8) Which of the following has dispersion forces as its only intermolecular force?
A) CH₄
B) HCl
C) C₆H₁₃NH₂
D) NaCl
E) CH₃Cl

9) Elemental iodine (I₂) is a solid at room temperature. What is the major attractive force that exists among different I₂ molecules in the solid?
A) London dispersion forces
B) dipole-dipole rejections
C) ionic-dipole interactions
D) covalent-ionic interactions
E) dipole-dipole attractions

10) The predominant intermolecular force in (CH₃)₂NH is __________.
A) London dispersion forces
B) ion-dipole forces
C) ionic bonding
D) dipole-dipole forces
E) hydrogen bonding

11) C₁₂H₂₆ molecules are held together by __________.
A) ion-ion interactions
B) hydrogen bonding
C) ion-dipole interactions
D) dipole-dipole interactions
E) dispersion forces

12) Which of the following has hydrogen bonding as its only intermolecular force?
A) HF
B) H₂O
C) C₆H₁₃NH₂
D) C₅H₁₁OH
E) None, all exhibit dispersion forces.

13) Which one of the following substances will have hydrogen bonding as one of its intermolecular forces?

14) Which one of the following substances will **not** have hydrogen bonding as one of its intermolecular forces?
15) The ease with which the charge distribution in a molecule can be distorted by an external electrical field is called the _________.
A) electronegativity
B) hydrogen bonding
C) polarizability
D) volatility
E) viscosity

16) ________ are particularly polarizable.
A) Small nonpolar molecules
B) Small polar molecules
C) Large nonpolar molecules
D) Large polar molecules
E) Large molecules, regardless of their polarity,

17) In which of the following molecules is hydrogen bonding likely to be the most significant component of the total intermolecular forces?
A) CH₄
B) C₅H₁₁OH
C) C₆H₁₃NH₂
D) CH₃OH
E) CO₂

18) Based on the following information, which compound has the strongest intermolecular forces?

<table>
<thead>
<tr>
<th>Substance</th>
<th>( \Delta H_{\text{vap}} ) (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Argon (Ar)</td>
<td>6.3</td>
</tr>
<tr>
<td>Benzene (C₆H₆)</td>
<td>31.0</td>
</tr>
<tr>
<td>Ethanol (C₂H₅OH)</td>
<td>39.3</td>
</tr>
<tr>
<td>Water (H₂O)</td>
<td>40.8</td>
</tr>
<tr>
<td>Methane (CH₄)</td>
<td>9.2</td>
</tr>
</tbody>
</table>

19) Based on molecular mass and dipole moment of the compounds in this table, which has the highest boiling point?

<table>
<thead>
<tr>
<th>Substance</th>
<th>Molecular Mass (amu)</th>
<th>Dipole Moment (D)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Propane, CH₃CH₂CH₃</td>
<td>44</td>
<td>0.1</td>
</tr>
<tr>
<td>Dimethylether, CH₃OCH₃</td>
<td>46</td>
<td>1.3</td>
</tr>
<tr>
<td>Methylchloride, CH₃Cl</td>
<td>50</td>
<td>1.9</td>
</tr>
<tr>
<td>Acetaldehyde, CH₃CHO</td>
<td>44</td>
<td>2.7</td>
</tr>
<tr>
<td>Acetonitrile, CH₃CN</td>
<td>41</td>
<td>3.9</td>
</tr>
</tbody>
</table>

20) Which one of the following should have the lowest boiling point?
A) PH₃
B) H₂S
C) HCl  
D) SiH₄  
E) H₂O

21) Of the following substances, __________ has the highest boiling point.  
A) H₂O  
B) CO₂  
C) CH₄  
D) Kr  
E) NH₃

22) Of the following, __________ has the highest boiling point.  
A) N₂  
B) Br₂  
C) H₂  
D) Cl₂  
E) O₂

23) Which one of the following derivatives of ethane has the highest boiling point?  
A) C₂Br₆  
B) C₂F₆  
C) C₂I₆  
D) C₂Cl₆  
E) C₂H₆

24) The intermolecular force(s) responsible for the fact that CH₄ has the lowest boiling point in the set CH₄, SiH₄, GeH₄, SnH₄ is/are __________.  
A) hydrogen bonding  
B) dipole-dipole interactions  
C) London dispersion forces  
D) mainly hydrogen bonding but also dipole-dipole interactions  
E) mainly London-dispersion forces but also dipole-dipole interactions

25) What intermolecular force is responsible for ice being less dense than liquid water?  
A) London dispersion forces  
B) dipole-dipole forces  
C) ion-dipole forces  
D) hydrogen bonding  
E) ionic bonding

26) What types of intermolecular forces exist between HI and H₂S?  
A) dipole-dipole and ion-dipole  
B) dispersion forces, dipole-dipole, and ion-dipole  
C) dispersion forces, hydrogen bonding, dipole-dipole, and ion-dipole  
D) dispersion forces and dipole-dipole  
E) dispersion forces, dipole-dipole, and ion-dipole

27) What type(s) of intermolecular forces exist between Br₂ and CCl₄?  
A) dispersion forces  
B) dispersion forces and ion-dipole  
C) dispersion forces and dipole-dipole
D) dispersion forces, ion-dipole, and dipole-dipole
E) None. Since both are gases at room temperature, they do not interact with each other.

II. Phase changes
28) Of the following, __________ is an exothermic process.
   A) melting
   B) subliming
   C) freezing
   D) boiling
   E) All are exothermic.

29) Which of the following is NOT a phase change?
   A) melting
   B) diffusion
   C) sublimation
   D) vaporization

30) The direct change of a substance from a solid to a gas is called _____.
   A) boiling
   B) evaporation
   C) sublimation
   D) condensation

31) The escape of gas molecules from the surface of an uncontained liquid is _____.
   A) evaporation
   B) condensation
   C) boiling
   D) sublimation

32) The first particles to evaporate from a liquid are _____.
   A) those with the lowest kinetic energy
   B) those farthest from the surface of the liquid
   C) those with the highest kinetic energy

33) Which of the following will evaporate the fastest?
   A) water at 20°C
   B) water at 40°C
   C) water at 0°C
   D) all of the above evaporate at the same rate.

III. Vapor Pressure
34) A volatile liquid is one that __________.
   A) is highly flammable
   B) is highly viscous
   C) is highly hydrogen-bonded
   D) is highly cohesive
   E) readily evaporates

35) Volatility and vapor pressure are __________.
   A) inversely proportional to one another
   B) directly proportional to one another
C) not related  
D) the same thing  
E) both independent of temperature

36) If a liquid is sealed in a container and kept at constant temperature, how does its vapor pressure change over time?  
A) It rises at first, then remains constant  
B) It rises at first, then falls  
C) It rises continuously

37) An increase in the temperature of a contained liquid _____.
A) has no effect on the kinetic energy of the liquid  
B) causes fewer particles to escape the surface of the liquid  
C) decreases the vapor pressure of the liquid  
D) causes the vapor pressure above the liquid to increase

38) When the vapor pressure of a liquid equals the atmospheric pressure, the liquid _____.
A) freezes  
B) boils  
C) condenses  
D) No change is observed

39) Water could be made to boil at 105°C instead of 100°C by _____.
A) increasing the air pressure on the water  
B) decreasing the air pressure above the water  
C) decreasing the pressure on the water  
D) applying a great deal of heat

40) Some things take longer to cook at high altitudes than at low altitudes because _____.
A) water boils at a lower temperature at high altitude than at low altitude  
B) water boils at a higher temperature at high altitude than at low altitude  
C) heat isn't conducted as well in low density air  
D) natural gas flames don't burn as hot at high altitudes  
E) there is higher moisture content in the air at high altitude

41) What is the pressure when a liquid is boiling at its normal boiling point?  
A) 505 kPa  
B) 101 kPa  
C) 202 kPa  
D) 0 kPa

42) Based on this figure, the boiling point of diethyl ether under an external pressure of 1.32 atm is _______ °C.  
A) 10  
B) 20  
C) 30  
D) 40  
E) 0

43) Based on this figure, the boiling point of
ethyl alcohol under an external pressure of 0.0724 atm is _____ °C.
A) 80  
B) 60  
C) 70  
D) 40  
E) 20

44) Based on this figure, the boiling point of water under an external pressure of 0.316 atm is ____ °C.
A) 70  
B) 40  
C) 60  
D) 80  
E) 90

IV. Phase diagrams

45) The phase diagram of a substance is shown to the right. The region that corresponds to the solid phase is _____.
A) w  
B) x  
C) y  
D) z  
E) x and y

46) On the phase diagram shown to the right, segment ________ corresponds to the conditions of temperature and pressure under which the solid and the gas of the substance are in equilibrium.
A) AB  
B) AC  
C) AD  
D) CD  
E) BC

47) On the phase diagram shown to the right, the coordinates of point ________ correspond to the critical temperature and pressure.
A) A  
B) B  
C) C  
D) D  
E) E

48) The normal boiling point of the substance with the phase diagram shown to the right is about _____°C.
A) 10  
B) 20  
C) 30  
D) 40
49) In this phase diagram, the area labeled _____ indicates the gas phase for the
substance.
A) w
B) x
C) y
D) z
E) y and z

50) According to this phase diagram, the normal boiling point of this substance is
_____ °C.
A) -3
B) 10
C) 29
D) 38
E) 0

51) In this phase diagram, the substance is
a ________ at 25 °C and 1.0 atm.
A) solid
B) liquid
C) gas
D) supercritical fluid
E) crystal

52) On a phase diagram, the melting point is the same as the __________.
A) triple point
B) critical point
C) freezing point
D) boiling point
E) vapor-pressure curve

V. Liquids & solids
53) A gas is _________ and assumes __________ of its container whereas a liquid is
_________ and assumes __________ of its container.
A) compressible, the volume and shape, not compressible, the shape of a portion
B) compressible, the shape, not compressible, the volume and shape
C) compressible, the volume and shape, compressible, the volume
D) condensed, the volume and shape, condensed, the volume and shape
E) condensed, the shape, compressible, the volume and shape

54) In liquids, the attractive intermolecular forces are __________.
A) very weak compared with kinetic energies of the molecules
B) strong enough to hold molecules relatively close together
C) strong enough to keep the molecules confined to vibrating about their fixed lattice points
D) not strong enough to keep molecules from moving past each other
E) strong enough to hold molecules relatively close together but not strong enough to keep
molecules from moving past each other

55) Which of the following is **not** a type of solid?
A) ionic
B) molecular
C) supercritical
D) metallic
E) covalent-network

56) As a solid element melts, the atoms become __________ and they have __________ attraction for one another.
A) more separated, more
B) more separated, less
C) closer together, more
D) closer together, less
E) larger, greater

57) Together, liquids and solids constitute __________ phases of matter.
A) the compressible
B) the fluid
C) the condensed
D) all of the
E) the disordered

58) Which statement is true about liquids but **not** true about solids?
A) They flow and are highly ordered.
B) They are highly ordered and not compressible.
C) They flow and are compressible.
D) They assume both the volume and the shape of their containers.
E) They flow and are not compressible.

**ANSWERS**

<table>
<thead>
<tr>
<th>Part I:</th>
<th>Part IV:</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) B</td>
<td>45) A</td>
</tr>
<tr>
<td>2) B</td>
<td>46) B</td>
</tr>
<tr>
<td>3) C</td>
<td>47) B</td>
</tr>
<tr>
<td>4) D</td>
<td>48) D</td>
</tr>
<tr>
<td>5) D</td>
<td>49) C</td>
</tr>
<tr>
<td>6) A</td>
<td>50) C</td>
</tr>
<tr>
<td>7) D</td>
<td>51) B</td>
</tr>
<tr>
<td>8) A</td>
<td>52) C</td>
</tr>
<tr>
<td>9) A</td>
<td></td>
</tr>
<tr>
<td>10) E</td>
<td>53) A</td>
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<td>11) E</td>
<td>54) E</td>
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<td>12) E</td>
<td>55) C</td>
</tr>
<tr>
<td>13) D</td>
<td>56) B</td>
</tr>
<tr>
<td>14) A</td>
<td>57) C</td>
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<tr>
<td>16) E</td>
<td></td>
</tr>
<tr>
<td>17) D</td>
<td></td>
</tr>
</tbody>
</table>
IM forces, Liquids & Solids—Free Response Practice

1. Explain each of the following in terms of atomic and molecular structures and/or intermolecular forces.
   (a) SbCl$_3$ has measurable dipole moment, whereas SbCl$_5$ does not.
   (b) The normal boiling point of CCl$_4$ is $77^\circ$C, whereas that of CBr$_4$ is $190^\circ$C.

2. Explain the following in terms of the electronic structure and bonding of the compounds considered.
   (a) The SO$_2$ molecule has a dipole moment, whereas the CO$_2$ molecule has no dipole moment. Include the Lewis (electron-dot) structures in your explanation.
   (b) At ordinary conditions, HF (normal boiling point = 20$^\circ$C) is a liquid, whereas HCl (normal boiling point = -114$^\circ$C) is a gas.

3. Consider the molecules CF$_4$ and SF$_4$.
   (a) Draw the complete Lewis electron-dot structure for each molecule.
   (b) In terms of molecular geometry, account for the fact that the CF$_4$ molecule is nonpolar, whereas the SF$_4$ molecule is polar.
4. Use principles of molecular bonding to explain why \( \text{H}_2 \) and \( \text{O}_2 \) are gases at room temperature, while \( \text{H}_2\text{O} \) is a liquid at room temperature.

5. The phase diagram for a pure substance is shown below. Use this diagram and your knowledge about changes of phase to answer the following questions.

(a) What does point \( V \) represent? What characteristics are specific to the system only at point \( V \)?
(b) What does each point on the curve between \( V \) and \( W \) represent?
(c) Describe the changes that the system undergoes as the temperature slowly increases from \( X \) to \( Y \) to \( Z \) at 1.0 atmosphere.
(d) In a solid-liquid mixture of this substance, will the solid float or sink? Explain.

6. The normal boiling and freezing points of argon are 87.3 K and 84.0 K, respectively. The triple point is at 82.7 K and 0.68 atmosphere.
(a) Use the data above to draw a phase diagram for argon. Label the axes and label the regions in which the solid, liquid and gas phases are stable. On the phase diagram, show the position of the normal boiling point.
(b) Describe any changes that can be observed in a sample of solid argon when the temperature is increased from 40 K to 160 K at a constant pressure of 0.50 atmosphere.
(c) Describe any changes that can be observed in a sample of liquid argon when the pressure is reduced from 10 atmospheres to 1 atmosphere at a constant temperature of 100 K, which is well below the critical temperature.
(d) Does the liquid phase of argon have a density greater than, equal to, or less than the density of the solid phase? Explain your answer, using information given in the introduction to this question.

7. Account for each of the following observations about pairs of substances. In your answers, use appropriate principles of chemical bonding and/or intermolecular forces. In each part, you answer must include references to both substances.
(a) Even though \( \text{NH}_3 \) and \( \text{CH}_4 \) have similar molar masses, \( \text{NH}_3 \) has a much higher boiling point \((-33 \text{ °C})\) than \( \text{CH}_4 \) \((-164 \text{ °C})\).
(b) At 25 °C and 1.0 atm, ethane \( (\text{C}_2\text{H}_6) \) is a gas and hexane \( (\text{C}_6\text{H}_{14}) \) is a liquid.
(c) Si melts at a higher temperature \((1,410 \text{ °C})\) than \( \text{Cl}_2 \) \((-101 \text{ °C})\).
MgO melts at a much higher temperature \( (2,852 \, ^\circ C) \) than NaF \( (993 \, ^\circ C) \).

8. Use appropriate chemical principles to account for each of the following observations. In each part, your response must include specific information about both substances.
   (a) At 25 °C and 1 atm, F\(_2\) is a gas, whereas I\(_2\) is a solid.
   (b) The melting point of NaF is 993 °C, whereas the melting point of CsCl is 645 °C.
   (c) The shape of the ICl\(_4^-\) ion is a square planar, whereas the shape of the BF\(_4^-\) ion is a tetrahedral.
   (d) Ammonia, NH\(_3\), is very soluble in water, whereas phosphine, PH\(_3\), is only moderately soluble in water.

9. Using the information from the table, answer the following questions about organic compounds.

<table>
<thead>
<tr>
<th>Compound Name</th>
<th>Compound Formula</th>
<th>( \Delta H_{\text{vap}} ) (kJ mol(^{-1}))</th>
</tr>
</thead>
<tbody>
<tr>
<td>Propane</td>
<td>CH(_3)CH(_2)CH(_3)</td>
<td>19.0</td>
</tr>
<tr>
<td>Propanone</td>
<td>CH(_3)COCH(_3)</td>
<td>32.0</td>
</tr>
<tr>
<td>1-propanol</td>
<td>CH(_3)CH(_2)CH(_2)OH</td>
<td>47.3</td>
</tr>
</tbody>
</table>

   (a) For each pair of compounds below, explain why they do not have the same value for their standard heat of vaporization, \( \Delta H_{vap} \). (You must include specific information about both compounds in each pair.)
   i. Propane and propanone
   ii. Propanone and 1-propanol

10. Use appropriate chemical principles to account for each of the following observations. In each part, your response must include specific information about both substances.
   (a) At 25°C and 1 atm, F\(_2\) is a gas, whereas I\(_2\) is a solid.
   (b) The melting point of NaF is 993°C, whereas the melting point of CsCl is 645°C.
   (c) Ammonia, NH\(_3\), is very soluble in water, whereas phosphine, PH\(_3\), is only moderately soluble in water.
Free Response Answers

1)  
(a) SbCl₃ has a measurable dipole moment because it has a lone pair of electrons which causes a dipole - or - its dipoles do not cancel - or - it has a trigonal pyramidal structure - or - a clear diagram illustrating any of the above.  
(b) CBr₄ boils at a higher temperature than CCl₄ because it has stronger intermolecular forces (or van der Waal or dispersion). These stronger forces occur because CBr₄ is larger and/or has more electrons than CCl₄.

2)  
(a)  
\[ \text{\begin{tikzpicture} \draw (-40:1cm) -- (-20:1cm) -- (20:1cm) -- (40:1cm) -- cycle; \end{tikzpicture}} \]

There is a dipole moment between the oxygen and the sulfur in sulfur dioxide and a bond angle of 119°. This results in a net dipole in the molecule. While there is a dipole in the carbon-oxygen bond, the 180° bond angle cancels the dipole moment in the molecule.  
(b) Hydrogen bonding (dipole-dipole attraction) is much larger in HF than in HCl.

3)  
(a)  
\[ \text{\begin{tikzpicture} \draw (-40:1cm) -- (-20:1cm) -- (20:1cm) -- (40:1cm) -- cycle; \end{tikzpicture}} \]

(b) In the tetrahedral CF₄, the polar C-F bonds are cancelled out by the equiangular pull of the 4 bonds. With an expanded octet and trigonal bipyramidal structure, SF₄ has a pair of unbonded electrons at the center of the bipyramid, this gives a "seasaw" shape to the molecule and an uneven pull to the polar S-F bonds.

4)  
H₂ and O₂ are both nonpolar substances that experience only London dispersion forces, which are too weak to form the bonds required in order for a substance to be a liquid at room temperature. H₂O is a polar substance whose molecules form hydrogen bonds with each other. Hydrogen bonds are strong enough to form the bonds required in a liquid at room temperature.

5)  
(a) Triple point. All three states of the substance coexist (equilibrium); the solid and the liquid have identical vapor pressures.  
(b) Curve VW represents the equilibrium between the liquid and its vapor. Along this line the liquid will be boiling. The points represent the vapor pressure of the liquid as a function of temperature.  
(c) At point X the substance is a solid, as its temperature increases (at constant pressure), at point Y the solid is in equilibrium with its vapor and will sublime. From point Y to Z it exists only as a vapor.
(d) Sink. A positive slope of the solid-liquid line indicates that the solid is denser than its liquid and, therefore, will sink.

6)  
(a) ![Graph showing solid, liquid, and gas phases with temperature and pressure axes.

(b) The argon sublimes.
(c) The argon vaporizes.
(d) The liquid phase is less dense than the solid phase. Since the freezing point of argon is higher than the triple point temperature, the solid-liquid equilibrium line slopes to the right with increasing pressure. Thus, if a sample of liquid argon is compressed (pressure increased) at constant temperature, the liquid becomes a solid. Because increasing pressure favors the denser phase, solid argon must be the denser phase.

7)  
(a) NH₃ has a much higher boiling point compared to CH₄ because CH₄ has only London dispersion forces while NH₃, which is polar, has dipole-dipole forces and even more specifically hydrogen bonding, which is stronger than regular dipole-dipole. Since NH₃’s intermolecular forces are bigger than those of CH₄, the bonds of NH₃ take more kinetic energy to break them and thus have a higher boiling point.
(b) Both have only LDF but since hexane is a more complex molecule, it has more electrons and thus more polarization opportunities. Because of this, hexane has more chances of including dipole forces and thus the intermolecular forces between hexane molecules are higher and require more kinetic energy to break. Thus, at 25°C, there is enough KE to make ethane into a gas while hexane is a liquid.
(c) Si has a covalent network bonding structure while Cl₂ has only LDF forces covalent network is a very rigid bonding structure and it’s difficult to break compared to the easily broken LDF Cl₂ bonds, thus Si melts at a higher temperature than Cl₂.
(d) Coulomb’s law states that the higher the atoms change, the stronger the bond between the atoms. Mg has a +2 charge and O has a -2 charge, which are greater than that of +1 and -1 charges of Na and F respectively. Thus, breaking the ionic bond between Na and F is easier than between Mg and O, and thus MgO melts at a higher temperature than NaF.

8)  
(a) F₂ and I₂ are both nonpolar molecules. The only forces between the molecules are London dispersion forces. These forces are created by the attraction of temporary dipoles. Molecules with more total electrons can form stronger dipoles, so the forces
are stronger. An I\textsubscript{2} molecule has almost 6 times as many total electrons as a F\textsubscript{2} molecule. The forces between I\textsubscript{2} molecules are much stronger than those between F\textsubscript{2} molecules. At 25\textdegree C, F\textsubscript{2} molecules have enough energy to overcome these forces, so they form a gas but I\textsubscript{2} molecules cannot overcome the forces so they form a solid.

(b) The equation for lattice energy (Coulomb’s law) is \( E = k\frac{q_1q_2}{r} \)

In both NaF and CsCl, both charges are 1, but the atoms of Cs and Cl are farther apart than the atoms of Na and F. The larger distance means it takes less energy to separate the Cs and Cl ions, so CsCl melts at a lower temperature. Cs has a larger atomic radius than Na, and Cl has a larger atomic radius than F because Cs and Cl have more filled electron shells than Na and F.

(c) I in ICl\textsuperscript{4} has six total e\textsuperscript{-} clouds surrounding it. Two of these are unshared pairs. The clouds arrange themselves in an octahedral geometry, and the two unshared pairs, which take up more room than bonds, take up opposite corners, leaving the 4 Cl in a square. B in BF\textsuperscript{4} has only four total e\textsuperscript{-} clouds around it and no unshared pairs. The clouds spread out in a tetrahedral geometry.

(d) NH\textsubscript{3} and PH\textsubscript{3} are both polar. However, only NH\textsubscript{3} can form hydrogen bonds with water molecules. This makes it much more soluble in water. The bonds in PH\textsubscript{3} are less polar, and P is too large to hydrogen bond.

9)

i. Propane is a hydrocarbon and the only IMF acting is London dispersion force. Propanone with a C=O is polar molecule and has strong IMF in addition to London dispersion force. Though they are three carbon derivatives, propanone has larger polarizing power and stronger IMF than propane. The heat of vaporization of propanone is greater than propane.

ii. In propanol, the presence of O-H group enables hydrogen bonding also compared to propanone. This is a strong IMF and hence its heat of vaporization is larger than that of propanone.

10)

(a) Elemental Fluorine and Iodine has only London dispersion force as their IMF. Since I\textsubscript{2} has more electrons than Chlorine, it more polarizable than F\textsubscript{2}. The London dispersion force will be stronger in Iodine than Fluorine hence the melting point is higher in Iodine than Fluorine.

(b) The lattice energy of CsCl is smaller than that of NaF. Sodium and Cesium ions have +1 charge and F and Cl ions have -1 charge. The size of Cs\textsuperscript{+1} ion is larger than that of Na\textsuperscript{+1} ion and Cl\textsuperscript{-1} ion is larger than F\textsuperscript{-1} ion. Lattice energy is given by \( E=\frac{q_1q_2}{r^2} \), where \( r \) is the distance between the ions. \( R \) is greater in CsCl and hence it has lower lattice energy than NaF. The melting point ionic solids depend up on the lattice energy of the ionic solid.

(c) NH\textsubscript{3} and PH\textsubscript{3} have the same geometry and they both are polar molecules. In NH\textsubscript{3} there is Hydrogen bonding possible with water molecules due to the lone pair of electrons while in PH\textsubscript{3} it is not possible though it has lone pair of electrons. This makes NH\textsubscript{3} readily soluble in water and PH\textsubscript{3} being nonpolar it will be only moderately soluble in water.