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2016-2017

**WS 12.1 Properties of Liquids and Solids (Viscosity, Surface Tension)**

1. What are the properties of a liquid?

2. What are the properties of a solid?

3. What are intermolecular forces? How are they different from intramolecular forces (i.e. ionic or covalent bonds)?

4. Why are intermolecular attractions so strong in liquids and solids, but so weak in gases?

5. What is viscosity?

6. What is surface tension?

7. What is volatility?

8. What is dynamic equilibrium?

9. What is vapor pressure?

10. How is a normal boiling point different from boiling point?

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11. Using your knowledge of evaporation, explain why perspiration cools you on a warm, dry day.

12. What do we mean when we say a liquid is volatile (Use your dictionary if necessary). Would you expect a volatile liquid to have a high or low vapor pressure? Explain.

13. Could vapor pressure be measured in an open container? Why or why not?

14. At room temperature, the vapor pressure of acetone (in nail polish remover) is approximately 220 torr, whereas the vapor pressure of ethyl alcohol is approximately 60 torr at this same temperature. Which of these substances possesses the stronger intermolecular attractions?

15. If I want to prepare a pot of spaghetti, would it be quicker to boil water in Mount Washington, NH (the highest summit in the northeast), or in Death Valley (the lowest point in the western hemisphere). Explain your answer. And by the way, does this necessarily mean that my pasta will cook faster?

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**WS 12.2 Intermolecular Forces**

1. For each of the compounds listed, identify the type(s) of intermolecular forces at work.

a. NF3

b. BH3

c. S8

d. H2O

e. C6H14

f. CH3OH

2. What are hydrogen bonds? Why are they more important in compounds containing N- H, O-H, and F-H bonds?

3. What type (or types) of intermolecular attractive forces are found in the following:

HCl Ar CH4

HF

NO+CO2

H2S

SO2



4. Rank the following compounds from strongest to weakest intermolecular forces: CH4, HF, NH3, H2O

5. Rank the following compounds from strongest to weakest intermolecular forces: C3H8, C2H5OH, C8H18

6. In terms of the type of intermolecular force present, explain why the smallest noble gas, helium, has a much lower boiling point than the largest gas, radon.

7. Rank the following in order of increasing boiling point: C9H18, CH4, C2H6,

8. Rank the following from lowest to highest boiling point? C7H15OH, CH3OH, C2H5OH

9. Which of the following has the highest boiling point: C2H6 or C2H5OH? Why?

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10. Which of the following would you expect to be more viscous? C8H16 or C4H10?Why?

11. Which of the following would have a greater vapor pressure? C9H20 or C2H6? Why?

12. Which of the following would have a greater surface tension: water or methane (CH4)?

Why?

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**WS 12.3 Intermolecular Forces**

1. For each of the following, identify what type of intermolecular forces are present: HF, C2H4, C2H5OH, NH3, CO2, AlCl3, H2O

2. Classify each of the following as intramolecular or intermolecular:

i. Hydrogen bonding in liquid water

ii. The O-H covalent bond in methanol, CH3OH

iii. The bonds that cause gaseous Cl2 to become a liquid when cooled.

3. Explain the difference between a hydrogen bond and a covalent bond.

4. Rank the three types of intermolecular forces from weakest to strongest. Explain why there is a difference in strength between the different types.

5. Sulfur dioxide, SO2, and carbon dioxide,CO2, have similar formulas, yet the boiling point of sulfur dioxide, is almost 70 degrees higher than that of carbon dioxide. Explain.

6. Predict which liquid in each pair below will show the largest vapor pressure. Explain your choice in each case.

a. H2O vs. CH3OH

b. CH3OH or CH3CH2CH2CH2OH

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7. For each of the following pairs of molecules, predict which you believe would have the highest boiling point, then explain your choice:

a. H2O or CH4

b. C2H6 or C6H14

8. For each of the following pairs, predict which you believe would be more viscous.

Explain your choice:

a. C14H30 or C7H16

b. AlCl3 or PCl3

9. For each of the following, which would you expect to have a lower surface tension.

Explain your choice:

a. H2O or CH4

b. C6H12 or C6H13OH

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**WS 12.4 Intermolecular Forces**

1. True or false: Hydrogen bonding is a type of London dispersion force (LDF).

2. Which type of intermolecular force is stronger: LDF or dipole?

3. What causes London dispersion forces?

4. The boiling points of the noble gas elements are listed below. Comment on the trend you observe. Explain this trend using your understanding of intermolecular forces.

He -272 °C Ne –245.9 °C Ar -185.7 °C

Kr -152.3 °C Xe -107.1 °C Rn -61.8 °C



5. For each of the following pairs, predict which you believe would have the lowest vapor pressure, then explain your choice:

a. NH3 or C2H6

b. C6H14 or C6H13OH

6. For each of the following pairs, predict which would have a higher boiling point? Explain your choice.

a. PH3 or H2O

b. PH3 or AlCl3

7. Imagine you are given the formulas for two substances. What would you look for to determine which substance has a higher surface tension.

8. True or false? A substance with high intermolecular forces will have a low viscosity.

9. Which of the following is more likely to be a liquid at room temperature: C4H10 or

C4H9OH? Explain why.

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**WS 12.5 Evaporation and Boiling (Heat of Fusion & Vaporization)**

1. Why does the vapor pressure of a liquid increase with increasing temperature?

2. Why is the boiling temperature of water less than 100C at high altitudes?

3. Methanol has a normal boiling point of 65C. It is a liquid at conditions of 1 atm and

25C. A small beaker filled with methanol is placed under a bell jar, and the air is then pumped out. It is observed that under a vacuum the methanol boils readily at 25C.

Use the kinetic molecular theory and the concept of vapor pressure to account for the lowered boiling point.

4. How many joules would be needed to convert 3 moles of water to vapor at 1000C? **(122,100 J)**

5. When water freezes, heat energy is released. How many joules are released to the surroundings when 3.6 kg of water change from liquid to solid? **(-1.2 x 106 J)**

6. How many kilojoules of energy are needed to convert 375 grams of water from ice to liquid? **(125.4 kJ)**

7. The molar heat of vaporization of methane, CH4, is 8.19 kJ/mol. If 97 grams of methane is made to boil, how much heat must be supplied? **(49.7 kJ)**

8. What is the heat of vaporization for water in calories per gram? **(540.4 cal/g)**

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9. How much energy would be required to vaporize each of the following:

a. 5.00 mol water **(203.5 kJ)**

b. 45.0 g water **(102 kJ)**

10. Calculate the molar heat of vaporization of a substance given that 0.43 moles of the substance absorbs 36.5 kJ of energy when it is vaporized. **(85 kJ/mol)**

11. Given that a substance has a molar mass of 259.0 g/mol and 71.8 grams of the substance absorbs 4.307 kJ when it melts, calculate the molar heat of fusion for this substance. **(15.5 kJ/mol)**

12. How much energy is needed to melt 25.4 gram of I2? The heat of fusion of I2 is 61.7 J/g. **(1570 J)**

13. How much energy will be needed to melt 4.24 grams of Pd? The enthalpy (heat) of fusion of Pd is 162 J/g. **(687 J)**

14. Calculate the kilojoules necessary to convert 550 grams of ethanol (C2H5OH) from liquid to vapor. Hvap = 38.6 kJ/mol. **(460 kJ)**

15. How many kilojoules of energy are necessary to melt 35 grams of benzene (C6H6) if Hfus

= 9.92 kJ/mol? **(4.5 kJ)**

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**WS 12.6 Phase Changes**

1. Sketch a heating curve for a substance with a boiling point of 134°C and a melting point of –15°C. Then answer the following questions. Note: the scale of the graph is not important, you should focus on the placement of the melting and boiling points.

a. Explain what is happening during the flat portions of your graph. Be sure to mention energy and the movement of molecules in your explanation.

b. What is the freezing point for this substance? The vaporization point?

2. Explain, on a molecular level, why the temperature remains constant as heat is added to vaporize a liquid at its boiling point.

3. Your teacher has asked you to calculate the energy required to heat a piece of copper from room temperature to 2000°C. What further information do you need to solve this problem? (Hint: there are seven things you will need to know.)

4. A 185 gram sample of water is heated from -15C to 75C. Calculate the energy required to make this change. **(1.3 X 105 J)**

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5. 95 grams of water is cooled (and condenses) from 120C to 34C. Determine the energy released by this action. **(-2.4 X105 J)**

6. You have a sample of H2O with a mass of 23.0 grams at a temperature of -460C. How many kilojoules of energy are necessary heat this H2O to 109C? **(72 kJ)**

7. Calculate the energy released when 10.0 grams of steam at 150C are converted into ice at –20.0 C. **(-31,555 J)**

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**WS 12.7 More Phase Changes**

1. Aluminum has a melting point of 660 C and a boiling point of 1800 C. Sketch and label a heating curve for aluminum.

2. Using words instead of calculations, list the steps outlining how we would determine the amount of energy needed to heat 200 grams of water from –13°C to 120°C.

3. How much energy is required to convert 100.0 grams of water at 20.0C completely to steam at 100C? **(2.60 X 105 J)**

4. Determine the energy needed when 55.6 grams of water at 43.2 C is heated to 78.1C. **(8,120 J)**

5. Determine the energy required when cooling 456.2 grams of water at 89.2 C to a final temperature of –19.5C. **(-341 kJ)**

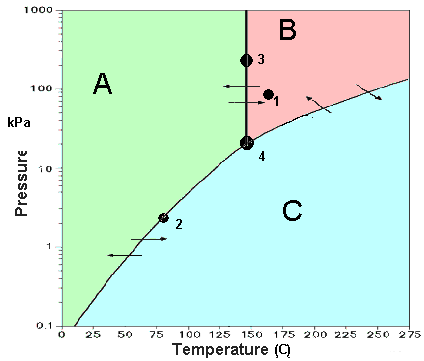
6. Determine the energy required to convert 32.1 grams of ice at –10.3C to liquid water at

0.0C. **(11,407 J)**

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**Worksheet 12.8**

1. What does any point on the curve between B & C represent?

2. What state(s) of matter is/are present at each of the following points?

a. 1 b. 2

3. What does point 4 represent?

c. 3 d. 4



4. Label each arrow on the graph with the appropriate phase change (boiling, melting, etc.)

5. Answer each of the following:

a. If I increase the temperature at point #2, holding the pressure constant, what phase change will occur?

b. If I decrease the pressure at point #1, holding the temperature constant, what phase change will occur?

c. If I decrease the temperature at point #1, holding the pressure constant, what phase change will occur?

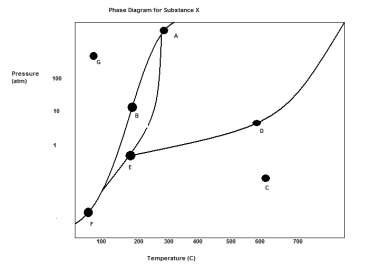
6. What is the normal boiling point of this substance? What is the normal melting point of this substance?

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**Worksheet 12.9**



Question 1: please label the following: solid, liquid, gas, sublimation, boiling, melting, condensation, freezing, deposition.



2. How is this graph different from the graph we have already observed?

3. Consider the substance represented by the graph above. What state(s) would be present at room temperature (assuming standard pressure)?

4. Explain how to find the normal boiling point and/or normal melting point of a substance. What is the normal melting point for this substance? What is its normal boiling point?

5. What are the pressure and temperature at triple point?

6. Is the critical point visible on this graph?

7. What state(s) would be present at the following points:

a.

b. f.

c. g.

d.

e.

8. If I am at point F and pressure is increased, while temperature is held constant,

what state of matter will be present? What process has occurred?

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