Chemistry H

Chapter 10
Study Guide – Answers

Name: _____

Date: _____ Per: _____

1. Describe the assumptions/postulates of the kinetic-molecular theory of gases:
   a. Gases are composed of tiny particles in constant rapid, ("random" or "straight–lined"?) motion.
   b. Gases are separated by relatively huge distances. The volume of the particles is essentially zero.
   c. The collisions between molecules are completely elastic. (No energy is converted to heat or other forms of energy.)
   d. The molecules of a gas display no attraction or repulsion.
   e. The average kinetic energy of the gas particles is directly proportional to the Kelvin temperature.
2. How do gas particles diffuse? Gas particles bounce off of one another as they mix transferring energy and spreading throughout a space filling it.
3. Fill in the following table.

Property	**Solid**	**Liquid**	**Gas**
Compressibility?	Low	Low	High
Density (high/low)	High	High	Low
Volume (variable/fixed)	Fixed	Fixed	Variable
Shape (variable/fixed)	Fixed	Variable	Variable
Diffusion (fast/slow)	Very Slow	Slow	Fast
Expansion (low/high)	Low	Low	High
Movement of Particles	Very Slow	Slow	Fast
Degree of Organization	Highly Organized	Random	Very Random
Intermolecular Forces	Holding tightly	Allowing movement	Broken

4. Temperature is best defined as: the average kinetic energy in a sample of matter
5. The state of a substance at a particular temperature depends mostly on: the strength of its intermolecular forces
6. List the characteristics of substances with strong intermolecular forces and weak intermolecular forces.

Strong Intermolecular Forces	Weak Intermolecular Forces
High melting point High surface tension High viscosity High melting point High heat of vaporization High heat of fusion More likely to form crystalline solids More likely to be liquids or solids at STP	Low melting point Low surface tension Low viscosity Low melting point Low heat of vaporization Low heat of fusion More likely to form amorphous solids More likely to be gases at STP

7. The fact that liquid water turns into water vapor when heated, rather than turning into hydrogen gas and oxygen gas indicates that intramolecular forces are stronger than intermolecular forces
8. What are the differences between a dipole and an induced dipole? A dipole is a polar molecule with permanent positive and negative ends. An induced dipole is non-polar and has only temporary positive and negative ends which can change charges from moment to moment.
9. Describe a dipole-dipole intermolecular force: the electrostatic force between the partial positive charge of one polar molecule and the partial negative charge of another polar molecule.

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10. Describe a hydrogen bond: a particularly strong dipole-dipole attraction between the hydrogen end of a polar molecule and the negative end (usually consisting of N, O, F) of another molecule
11. Describe an induced dipole – induced dipole intermolecular force: the weak electrostatic force between two non-polar molecules that occurs as electrons shift within the molecules creating temporary partial positive and negative charges
12. What determines the strength of an induced dipole – induced dipole force? The total number of electrons in the substance. The greater the number of electrons, the greater the potential shift of the cloud and the greater the partial charge.
13. Describe the concept of viscosity: the thickness or cohesiveness of a substance. The greater the intermolecular force of a substance, the greater the viscosity. Viscosity increases as temperature decreases.
14. Name some of the unique properties of water. What accounts for these unique properties?
    Properties: high boiling point, universal solvent, high specific heat, high surface tension, high heat of vaporization
    Cause: mostly these characteristics are caused by the strong intermolecular forces (hydrogen bonds) between water molecules
15. Describe a unit cell: the smallest repeating geometric unit that occurs in a crystalline substance
16. Describe a hydrate: a crystalline structure in which water molecules are attached or trapped
17. Describe each of the 4 types of solid substances:
    Metallic: frameworks of metal atoms in which the electrons flow freely between atoms in a "sea of electrons"
    Covalent: substances in which molecules are held together by intermolecular forces such as dipole-dipole, dispersion forces or hydrogen bonds
    Ionic: repeating geometric patterns of positively and negatively charged ions held in place by electrostatic attractions
    Covalent Network: Substances in which atoms are held in a crystalline structure by covalent bonds
18. Changing the temperature of a substance involves changing the substance's: kinetic energy
19. Changing a phase of a substance involves changing the substance's: potential energy
20. Classify each of the phase changes as either endothermic or exothermic.

Phase Change	**Energy Change**
Sublimation	Endothermic – separating particles stores energy
Melting	Endothermic – separating particles stores energy
Vaporization	Endothermic – separating particles stores energy
Deposition	Exothermic – particles connecting releases energy
Freezing	Exothermic – particles connecting releases energy
Condensation	Exothermic – particles connecting releases energy

21. Describe the flat portions of a heating curve: the flat portions represent changes in potential energy
22. Describe the sloped portions of a heating curve: the sloped portions represent changes in kinetic energy

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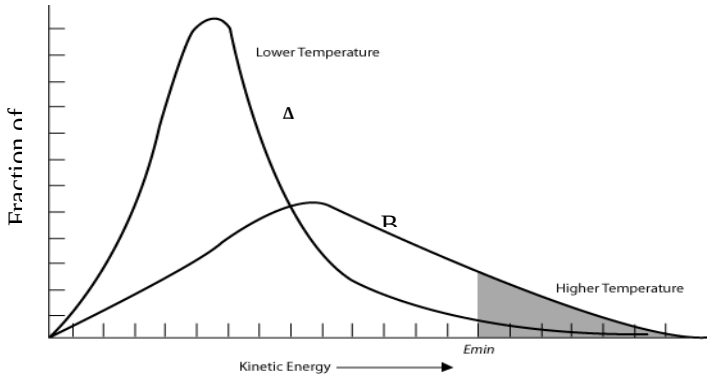
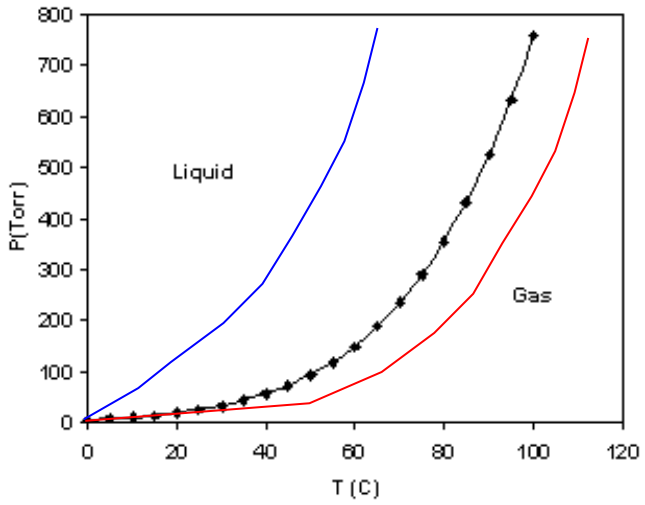


23. Which sample of gas has more particles ready to evaporate? Line b has more particles beyond the E_{min} point
24. Which distribution has the lower average energy? Line A has the lowest average kinetic energy. The taller peak only means that Line A has a large portion of its molecules at a lower temperature.



25. What is the vapor pressure of the substance above at 60°C? Approximately 150 torr
26. At what temperature would this substance boil if the atmospheric pressure was 500Torr? Approximately 90°C
27. Draw a line on the graph above showing a substance with higher intermolecular forces than those of the substance shown. The red line has greater intermolecular forces.
28. Draw a line on the graph above showing a substance with lower intermolecular forces than those of the substance shown. The blue line has lower intermolecular force.

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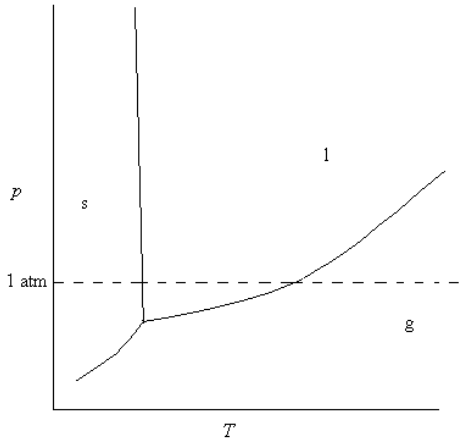

29. What do the lines on the graph above represent? Each line represents an equilibrium position between two phases at a particular temperature and pressure. Equilibrium exists anywhere along the line.
30. Label each of the six phase changes on the graph. Check the text.
31. Label the triple point on the graph. Check the text or notes.
32. How can a solid be converted into a gas without changing its temperature? By reducing the pressure
33. How can a gas be converted into a liquid without changing its pressure? By lowering its temperature
34. Calculate the amount of heat in joules needed to melt 120.0 g of ice at 0 ºC. ($\Delta H_{fus(H2O)}$ = 6.02 kJ/mol)

$$\frac{120.0\ \cancel{g\ H_2O}}{} \left| \frac{1\ \cancel{mol\ H_2O}}{18.015\ \cancel{g\ H_2O}} \right| \frac{6.02\ \cancel{kJ}}{1\ \cancel{mol\ H_2O}} \left| \frac{1000\ J}{1\ \cancel{kJ}} \right. = 40099.9\ J \quad \boxed{40100\ J}$$

35. Calculate the energy absorbed when 195 g of dry ice (CO_2) sublimes. ($\Delta H_{sub(CO2)}$ = 196.3 J/g)

$$\frac{195\ \cancel{g\ CO_2}}{} \left| \frac{196.3\ J}{1\ \cancel{g\ CO_2}} \right. = 38278\ J \quad \boxed{38300\ J}$$

36. What mass of aluminum metal would absorb 430.0 kJ when melted at its melting point? ($\Delta H_{fus(Al)}$ = 396.6 J/g)

$$\frac{430.0\ \cancel{kJ}}{} \left| \frac{1000\ \cancel{J}}{1\ \cancel{kJ}} \right| \frac{1\ g\ Al}{396.6\ \cancel{J}} = 1084.2\ g\ Al \quad \boxed{1084\ g\ Al}$$

37. Calculate the amount of heat in joules needed to vaporize 130.0 g of liquid water at its boiling point. (The molar enthalpy of vaporization of water is 40.7 kJ.)

$$\frac{100.0\ \cancel{g\ H_2O}}{} \left| \frac{1\ \cancel{mol\ H_2O}}{18.015\ \cancel{g\ H_2O}} \right| \frac{40.7\ \cancel{kJ}}{1\ mol\ \cancel{H_2O}} \left| \frac{1000\ J}{1\ \cancel{kJ}} \right. = 225922\ J \quad \boxed{226000\ J}$$

38. How many joules of energy are required to melt a 10.0 lb bag of ice at 0°C? ($\Delta H_{fus(H2O)}$ = 6.02 kJ/mol & 1 lb = 2.205 kg)

$$\frac{10.0\ \cancel{lb\ H_2O}}{} \left| \frac{2.205\ \cancel{kg\ H_2O}}{1\ \cancel{lb\ H_2O}} \right| \frac{1000\ \cancel{g\ H_2O}}{1\ \cancel{kg\ H_2O}} \left| \frac{1\ \cancel{mol\ H_2O}}{18.015\ \cancel{g\ H_2O}} \right| \frac{6.02\ \cancel{kJ}}{1\ \cancel{mol\ H_2O}} \left| \frac{1000\ J}{1\ \cancel{kJ}} \right. = 7368359\ J \quad \boxed{7.37 \times 10^6\ J}$$