Name ________________________________**Chem 162, Section**: _______**Group Number:** ______

ALE 17. Phase Changes and Phase Diagrams

(Reference: 12.1 – 12.2 *Silberberg 5th edition*)

How do temperature and pressure determine the phase of matter?

The Model: Heat and Phase Changes

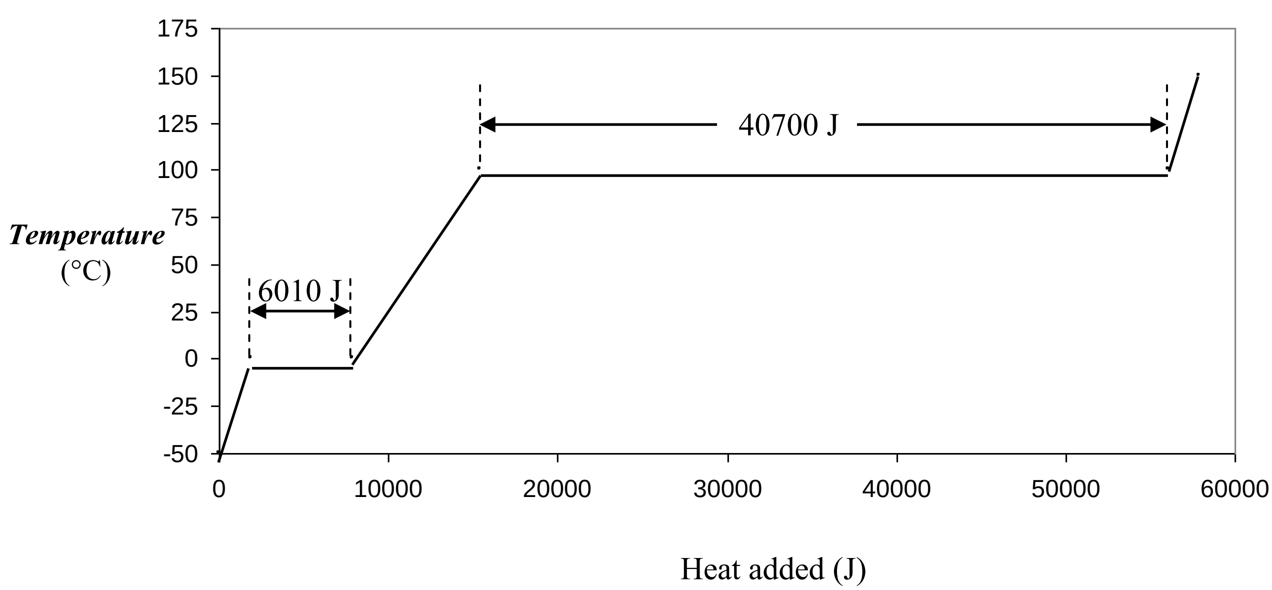
The three phases of a substance that are of interest to chemists are solid, liquid, and gas. Depending on the pressure that the substance is experiencing, a temperature can be found that would put two phases in **equilibrium** with each other. Equilibrium is established when the number of molecules in each phase is no longer changing with time. The equilibrium is **dynamic**. That is, a molecule at one location of the system may go from phase α to phase β, but simultaneously somewhere else in the system another molecule goes from phase β to phase α.

at 0.00 °C and 1 atm: $H_2O(s) \underset{\text{crystallization}}{\overset{\text{fusion}}{\rightleftharpoons}} H_2O(l)$ $\Delta H^{o}_{fus} = 6.01$ kJ

at 100.00 °C and 1 atm: $H_2O(l) \underset{\text{condensation}}{\overset{\text{vaporization}}{\rightleftharpoons}} H_2O(g)$ $\Delta H^{o}_{vap} = 40.7$ kJ

at 0.00 °C and 1 atm: $H_2O(s) \xrightarrow{\text{sublimation}} H_2O(g)$ $\Delta H^{o}_{subl} = 46.7$ kJ

A pure substance melts at a single temperature, the freezing point (t_f). If the ambient pressure is 1 atm, it is called the **normal freezing point** (and symbolized as t_f°). The following **heating curve** follows the temperature of 1 mole of water as heat is added to it. (The amount of added heat is zero when the water starts as ice at -50 °C.)



Key Questions

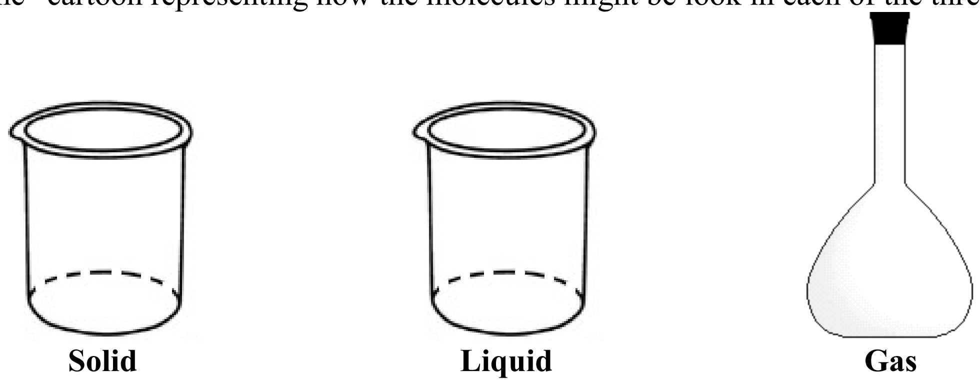
1. Label the heating curve indicating when the water exists as: ① a solid; ② an equilibrium of solid and liquid; ③ a liquid; ④ an equilibrium of liquid and gas; and ⑤ a gas.
2. What is an "everyday word" that may be used instead of:

a. "fuses" to describe a solid becoming a liquid? ______________________

b. "crystallizes"? ___________________________

c. "vaporizes"? ___________________________

3a. In each of the "containers" below, using twelve ●'s to represent 12 water molecules, sketch a "freeze frame" cartoon representing how the molecules might be look in each of the three phases.



b. What differentiates solid, liquid and gas phases from each other? As part of your response discuss how the energy of attraction between particles compares with their energy of motion (i.e. their kinetic energy, E_k), the distances between particles, their compressibility and their relative viscosities (i.e. resistance to flow).

c. What types of forces, intramolecular or intermolecular…..

i.) prevent ice cubes from adopting the shape of their container? ___________________________

ii.) are overcome when ice melts? ___________________________

iii.) are overcome when liquid water is vaporized? ___________________________

iv.) are overcome when gaseous water is converted to $H_{2(g)}$ and $O_{2(g)}$? ______________________

d. How is one phase converted into another phase? *Explain*.

4a. According to the heating curve on page 1, which one is true? (Circle your answer.)

i. The temperature decreases as more and more of the liquid is converted into a gas.

ii. A liquid boils at a single temperature.

iii. The temperature increases as more and more of the liquid is converted into a gas.

b. Explain why this observation is reasonable. (Where does the heat energy "go" during a phase change? What happens to the temperature as heat is added during a phase change? Why?)

c. Why is the heat of fusion (ΔH_{fus}) of a substance smaller than its heat of vaporization (ΔH_{vap})?

d. Why is the heat of sublimation (ΔH_{subl}) of a substance greater than its ΔH_{vap}?

e. At a given temperature and pressure, how does the magnitude of the ΔH_{vap} of a substance compare with that of its heat of condensation? Explain.

Model: Calculating the Amount of Heat Involved in Phase Changes

Recall the equation: $q = C \cdot m \cdot \Delta T$ (where q is the heat energy, C is the specific heat capacity of the substance undergoing a temperature change, m is the mass of the substance, and ΔT is the change in the temperature: $\Delta T = t_f - t_i$). This equation works ONLY when there is **no phase change**. You use it to find how much energy is required to change the temperature of a certain substance as long as you know the specific heat of the substance and as long as no change in state occurs.

Now for when there is a change in state: The energy required for a change of state is given a special name called enthalpy. So the "enthalpy of vaporization" (symbolized by ΔH_{vap}) of water is the energy needed to vaporize (or boil) one gram of water when the water is already at its boiling point. The ***ΔH_{vap} of water*** is 2260 J/g (or 40.7 kJ/mol). So for each gram of *liquid water* at 100.0^{o}C, 2260 J are required to vaporize it. Similarly, the enthalpy of fusion (**ΔH_{fus}**) **of H_2O** is 334 J/g (or **6.01 kJ/mol**). So each gram of ice at 0.0^{o}C there are 334 J of energy required to melt it. (Fusion is melting.)

Key Questions

5a. How much heat (in J) must be added to 1 mol of liquid water at 100 °C to convert it into steam?

b. Suppose you have 1.00 mol (*i.e.*, 18.02 g) of ***liquid*** water at 0.0 °C and add just enough heat to bring the temperature of the water to 100.0 °C, but no more heat is added. How much heat was added? What fraction of the water is converted into a gas? *Show your work using units and sig. figs.*

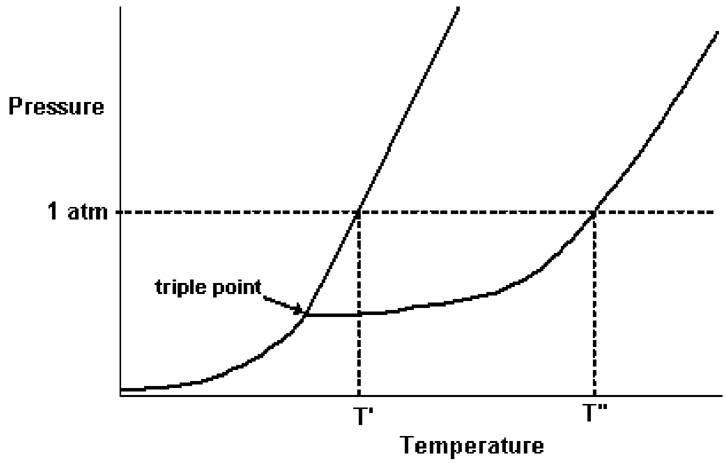
c. Suppose you have 1.00 mol of liquid water at 0.0 °C and add 7540 J of heat to it (in order to get its temperature to 100 °C) and then you add an additional 8140 J of heat to it. What fraction of the water will have been converted into steam? At what temperature will the system be? *Show your work.*

6. Use the data below to calculate the total heat in Joules needed to convert 0.333 mol of ethanol gas at 300.^{o}C and 1 atm to liquid ethanol at 25.0 ^{o}C and 1 atm. Clearly show your work using dimensional analysis and correct significant figures.

BP of ethanol @ 1.00 atm = 78.5 ^{o}C; ΔH^o_{vap}= 40.5 kJ/mol; c_{gas} = 1.43 J/g ^{o}C; c_{liq} = 2.45 J/g ^{o}C

The Model: Phase Diagrams

A **phase diagram** shows you the combinations of pressure and temperature that will result in a substance being a solid, a liquid, or a gas. The following is a sketch of a phase diagram of the typical substance (but not of water). The thick curves represent the combinations of pressure and temperature that result in an equilibrium between two (or perhaps all three) of the phases. The dashed lines are drawn to emphasize a temperature that results in an equilibrium between two phases when the ambient pressure is 1 atm.



Key Questions

7. Label the phase diagram above with "solid", "liquid", and "gas" in the appropriate regions.
8. Imagine that you have a substance in a cylinder and that there is a piston that you could push in or pull out to adjust the pressure that the substance is experiencing. You can also place the cylinder with the substance in a thermal bath to adjust its temperature.

a. What combination of pressure and temperature is likely to result in a solid? (Circle your choice.)

i. low pressure, low temperature
ii. low pressure, high temperature
iii. high pressure, low temperature
iv. high pressure, high temperature

b. What combination of pressure and temperature is likely to result in a gas? (Circle your choice.)

i. low pressure, low temperature
ii. low pressure, high temperature
iii. high pressure, low temperature
iv. high pressure, high temperature

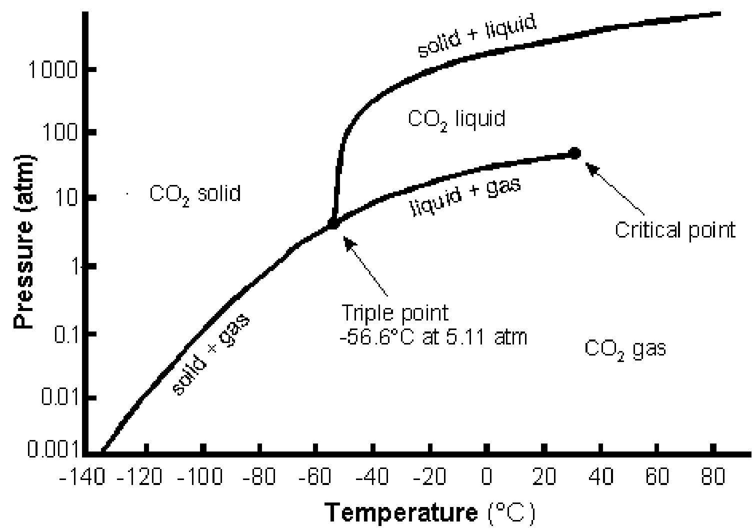
9. What is the significance of the "triple point"? (*i.e.*, What makes it special?)

10. What are the temperatures T′ and T″ called?

T′ = ______________________________

T″ = ______________________________

11. Below is the phase diagram of carbon dioxide. (Note the exponential scale on the *y*-axis.)



a. Solid CO_2 is referred to as "dry ice". Explain why dry ice *sublimes* at room conditions (i.e. 1 atm) and does not melt. *Use the phase diagram of CO_2 to illustrate your explanation.*

b. At what temperature does dry ice sublime at a pressure of 760 mmHg?_____________
Show how you arrived at your answer on the phase diagram above.

c. Suppose a sample of CO_2 is placed in a cylinder with a piston and maintained at -40 °C. If the pressure of the CO_2 started at 1 atm and increased until it reached 1000 atm, what would happen to the CO_2? At what pressure(s) would this(these) change(s) occur? *Show how you arrived at your answer on the phase diagram above*.

d. The "critical temperature" is the temperature above which a substance cannot be liquefied, no matter how much pressure it is under. What is T_c for CO_2? __________
Show how you arrived at your answer on the phase diagram above.

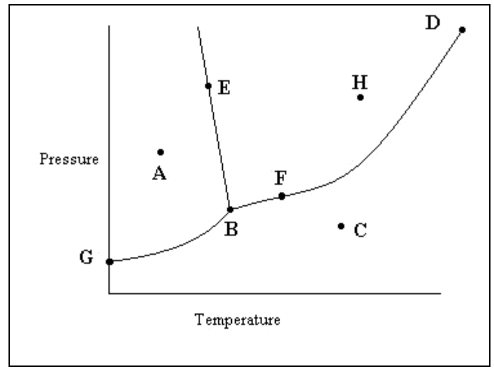
12. Consider the phase diagram, below, for substance "X".

    a.) What phase(s) is (are) present at the following points?

    A = E =

    B = F =

    C = H =



    b.) Which point corresponds to the critical point?

    c.) Which curve corresponds to the conditions at which the solid-gas equilibrium exists?

    d.) Describe what happens when you start at point A and increase the temperature a constant pressure.

    e.) Describe what happens when you start at point H and decrease the pressure a constant temperature.

    f.) Is liquid "X" more or less dense than solid "X"? Explain your reasoning.