W/S Phase Changes

1. A phase is a homogeneous part of the system in contact with other parts of the system but separated from them by a well-defined boundary. Phase changes are transformations from one phase to another (occurs when nrg is added or removed from a substance).

2. Equilibrium vapor pressure is the vapor pressure (force exerted by the vapors of a liquid on the surface of that liquid in a closed container) when a dynamic equilibrium exists between condensation & evaporation (rate at which the liq. is condensing & evaporating is equal). As temp increases, vapor pressure increases.

Figure 11.3 pg 496 talks about how its measured.
3. Dynamic equilibrium is reached when the rate (speed) of a forward process (like vaporization) is exactly balanced by the rate of the reverse process (condensation).

4. Molar heat of vaporization ($\Delta H_{\text{vap}}$) - energy required to vaporize 1 mole of a liquid. Units: kJ/mol

Molar heat of fusion ($\Delta H_{\text{fus}}$) - energy required to melt 1 mole of a solid. Units: kJ/mol

Molar heat of sublimation ($\Delta H_{\text{sub}}$) - energy needed to sublime 1 mole of a solid. Units: kJ/mol

\[ \Delta H_{\text{sub}} = \Delta H_{\text{vap}} + \Delta H_{\text{fus}} \]

5. $\Delta H_{\text{sub}} = \Delta H_{\text{vap}} + \Delta H_{\text{fus}}$

Based on Hess's law

6. The molar heat of vap. ($\Delta H_{\text{vap}}$) is directly related to strength of IMF's that exist in a liq. If the IMF's are strong, it takes a lot of energy to free the molecule from the liq. phase and the $\Delta H_{\text{vap}}$ will be high. Such liq will also have a low vapor pressure.

7. Boiling point is the temp @ which the vapor pressure of a liquid is equal to the external pressure. The higher the external pressure, the higher the BP.
8. \[ \frac{74.60 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \times \frac{40.79 \text{ kJ}}{18.02 \text{ g H}_2\text{O}} = 169 \text{ kJ} \]

9. \[ q_1 = 866.6 \text{ g log (2.03 J/g °C) (0°C - -10°C)} \]
   \[ q_1 = 17.6 \text{ kJ} \]

10. \[ q_{10} = 866.6 \text{ g log (1.99 J/g °C) (124°C - 100°C)} \]
    \[ q_{10} = 289 \text{ kJ} \]

11. \[ q_{11} = 68.09 \text{ g (1.99 J/g °C) (100°C - 124°C)} \]
    \[ q_{11} = 3.25 \text{ kJ released} \]

12. \[ q_2 = 68.09 \text{ g (1 mol) - 40.79 kJ} \]
    \[ q_2 = -154 \text{ kJ} \]
\[ q_3 = 68.0 \times (4.184 \text{ J/g} \cdot ^\circ\text{C}) (45 ^\circ\text{C} - 100 ^\circ\text{C}) \]
\[ q_3 = -15.16 \text{ kJ} \]

**Total heat released** = \[ 3.25 + 154 + 15.6 \]

\[ *\text{no need for neg sign} \]
\[ = 173 \text{ kJ} \]

**11.**

a) Other factors being equal, liquids evaporate faster at higher temps.

b) The greater the surface area, the greater the rate of evaporation.

c) Weak IMF’s imply a high vapor pressure & rapid evaporation.

**12.**

\[ \Delta H_{\text{vap}} = \Delta H_{\text{sub}} - \Delta H_{\text{fus}} = (62.30 \text{ kJ/mol}) - (15.27 \text{ kJ/mol}) \]

\[ = 47.03 \text{ kJ/mol} \]

**13.** The substance w/ the lowest boiling point will have the highest vapor pressure at some particular temp. Thus, butane will have the highest vapor pressure (at \(-10 ^\circ\text{C}\)) & toluene the lowest.
b) The solid-liquid boundary line has a negative slope. The neg. slope is due to the fact that the molar volume of ice (volume of 1 mol of ice vs. the volume of 1 mol of water) is greater than that of liq. water; hence water is denser than ice.

C) i) raising the temp @ constant pressure. Starting @ A is in the solid phase. The solid would warm until melting occurs, if warming continues, the liq. water would then boil, if warming continued the temp of the steam would inc.

ii) @ point C water is in the gas phase, cooling w/out changing pressure would eventually result in the formation of ice. Liquid water would never form.

iii) @ B water is in the liq. phase. Lowering pressure w/out changing the temp would eventually result in boiling & conversion to vapor.
15) a) 3  b) under atmospheric conditions the rhombic allotrope is more stable.
   c) @ 80°C & 1 atm the stable allotrope is rhombic sulfur, as temp is inc. there is first a transition to the monoclinic allotrope & as temp is inc. further the solid melts.

16) B

17) a) Solid  b) vapor

18) *assuming that liquid remains in the container after equilibrium has been established

19) a) dec  b) no change  c) no change

20) \[ \text{CaCO}_3(s) \rightarrow \text{CaO}(s) + \text{CO}_2(g) \]

   3 phases; CaCO\(_3\) & CaO constitute 2 (2 solid & 1 gas) separate phases b/c they are separated by well defined boundaries.

21) a) 2 (diamond/graphite/liquid
    & graphite/liquid/vapor)

   b) diamond  c) high pressure & temp
2 main reasons:

1. As water freezes, heat is released \( \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(s) \quad \Delta H_{\text{fus}} = -6.01 \text{ kJ/mol} \)

   the heat released protects the fruit.

   Of course spraying the trees w/ warm water is even more helpful.

2. The ice forms an insulating layer to protect the fruit.

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(23) \[
\begin{align*}
\text{Pressure (atm)} & \quad \text{Temperature (°C)} \\
0 & \quad -40 \\
10 & \quad 0 \\
20 & \quad 40 \\
30 & \quad 80 \\
40 & \quad 120
\end{align*}
\]

*Normal boiling point = temp @ which a liquid boils when external pressure is equal to 1 atm

*Critical temp = highest temp @ which a substance can exist as a liquid

*Critical pressure = min pressure that must be applied to bring about liquification(@ Tc)
plotting the normal & boiling point & critical temp/pressure as two points on the graph & connecting the two w/a straight line gives us an idea of where the liquid-vapor phase boundary exists for CCl₄F₂. (In reality this phase boundary is curved)

We see the point in question (20°C, 18 atm) lies on the liquid side of the phase boundary, so yes if the gas is compressed to this point it will condense to a liq.