SOLIDS, LIQUIDS & PHASE CHANGES

Chang, pp. 513-514 # 11.59 - 11.86

11.59 A phase change occurs when a material changes from one physical state to another.

solid → liquid (melt) liquid → solid (freeze)
liquid → gas (boil) gas → liquid (condensation)
solid → gas (sublimation) gas → solid (reverse sublimation or deposition)

11.60 Force per unit area exerted on the surface of a liquid by the liquid's own vapor, once vapor equilibrium has been established. It is measured with a manometer. Vapor pressure increases with temperature increase.

11.61 (Opposing processes balance each other out. For example if evaporation and condensation happen at the same rate the system is at equilibrium.)

11.62 ⓐ molar heat of vaporization - the amount of energy (at its B.P.) needed to vaporize (change from liquid to gas) one mole of liquid.
ⓑ molar heat of fusion - the amount of energy needed to change one mole of solid at its M.P. to a liquid.
ⓒ molar heat of sublimation - the amount of energy needed to change one mole of solid at its subl. point to a gas.
Units for all of these are typically KJ/mol (or J/mol)

11.63 $\Delta H_{sub} = \Delta H_{fus} + \Delta H_{vap}$ (Hess's Law)

11.64 The greater the ΔH_{vap}, the stronger the IMF's

11.65 False - the opposite is true. A greater ΔH_{vap} means the molecules are held together more tightly, therefore fewer of them will become vapor at a given temp.

11.66 BP (Def. 1) - temp. at which a liquid may change to a gas.

11.66 (cont) (Def 2) - the temp. at which the vapor pressure of a liquid equals atmospheric pressure.
External pressure (along with IMF's) helps hold a substances molecules in the liquid state, so the greater the atmospheric pressure, the more difficult it is for molecules to break away from the liquid, thus they need more energy (a higher temperature) to boil. 65°C

11.67 As boiling progresses, energy which is being put into the system is used to break apart the attractions which hold particles in the liquid state. Prior to that, the energy was used to accelerate molecular motion (and after!)

11.68 Critical Temperature is the highest temperature at which a gas can be liquefied by increasing the pressure on it.

11.69 The stronger the IMF's, the higher the B.P. and the higher the C.T. Water has hydrogen bonding, which is the strongest of IMF's, making water's B.P. and C.T. higher than other substances with similar sized molecules.

11.70 Boiling points increase as pressure increases for both substances. For carbon tetrachloride (as with most substances) the melting point also increases with pressure. Water is unique, because its hydrogen bonding leads to a crystalline pattern with "holes" between the molecules. Increasing pressure disrupts this pattern, turning ice into liquid. Therefore, water's M.P. decreases as pressure increases.

11.71 Solid CO_2 is called dry ice because it looks like ice, but does not melt. Under atmospheric pressure, it sublimes.

11.72 The higher the humidity, the more likely vapor from the air is to be condensed back on to the clothes. As clothes dry, the water molecules which evaporate from the clothes have the ever present possibility of returning to the clothes (along with any vapor which was already in the air), keeping them wet or damp. The less moisture there is in the air, the less likely this is to happen. To consider the extreme, would clothes dry more quickly on a 60°F, clear, cloudless day or an 80°F day with pouring rain? (Additional note: temperature DOES play a role, since a higher percentage of molecules with attain the minimum energy "threshold" to evaporate at a higher temp.)

11.73 (a) $\Delta H_{vap}(H_2O) > \Delta H_{fus}(H_2O)$ (almost 7 times as much energy must be put in to boil H_2O than that needed to melt ice.)

(40.79 KJ/mol) (6.01 KJ/mol)

11.74 No. It does not matter where the energy comes from, you could use a nuclear reactor and water will still boil at 100°C (if pressure is = standard atmospheric)

11.75 $q = m\Delta H_v = n\,\Delta H_{v\,(\text{molar})}$

$= (74.6\cancel{g})\left(\frac{1\cancel{mol}}{18.02\cancel{g}}\right)(40.79\ KJ/\cancel{mol}) = 168.8642619\ KJ$

$\approx \boxed{169\ KJ}$

11.76 ① $q = mc\Delta t$

warm ice to M.P. → $= (866g)(2.03\ J/gC°)(0°C - (-10°C)) = 17579.8\ J \rightarrow 17.5798\ KJ$

melt ice → ② $q = m\Delta H_f$ or $n\,\Delta H_{f\,(\text{molar})}$

$= (866g)\left(\frac{1mol}{18.02g}\right)(6.01\ KJ/mol) = 288.826859\ KJ$

warm meltwater to B.P. → ③ $q = mc\Delta t$

$= (866g)(4.184\ J/gC°)(100°C - 0°C) = 362334.4\ J \rightarrow 362.3344\ KJ$

boil → ④ $q = m\Delta H_v$ or $n \cdot \Delta H_{v\,(\text{molar})}$

$= (866\cancel{g})\left(\frac{1\cancel{mol}}{18.02\cancel{g}}\right)(40.79\ KJ/\cancel{mol}) = 1960.2744\ KJ$

heat the steam to 126°C → ⑤ $q = mc\Delta t$

$= (866g)(1.99\ J/gC°)(126°C - 100°C) = 44806.84\ J \rightarrow 44.80684\ KJ$

TOTAL = 2673.822039 KJ

$\boxed{2670\ KJ}$

11.77 ⓐ The higher the temperature, the faster the evaporation rate.

ⓑ The greater the surface area (more spread out) the substance has, the faster the evaporation rate.

ⓒ The stronger the IMF's, the slower the evap. rate.

11.78 ~~solid~~ 15.27 KJ/mol { Solid ↓ Liquid

62.30 KJ/mol − 15.27 KJ/mol { ↓ Gas

$\Delta H_{\text{subl.}} = 62.30\ KJ/mol$

$\Delta H_{vap} = \boxed{47.03\ KJ/mol}$

11.79 - Lowest B.P. = weakest IMF's = highest P_{vap}
so butane would have the lowest B.P. & highest P_{vap}
while toluene would have the highest B.P. & lowest P_{vap}

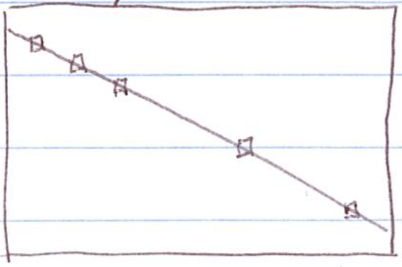
11.80 - Freezing, then sublimation

11.81 - Freezing, then evaporation of ice

11.82 - Steam has 40.79 KJ/mol more energy
than liquid water at the same temp. (100°C).

11.83

t(°C)	P(torr)	$1/T(K)$ ⓧ	lnP Ⓨ
200	17.3	0.06211	2.8507
250	74.4	0.00191	4.3095
300	246.8	0.00175	5.5086
320	376.3	0.00169	5.9304
340	557.9	0.00163	6.3242



$$\ln P = -7194.29934\left(\frac{1}{T}\right) + 18.1$$

$$\Delta H_{vap} = -R \cdot \text{Slope}$$
$$= (-8.31\ J/mol\ K) \cdot (-7194.29934\ K) = 59784.62752$$
$$= \boxed{59.8\ KJ/mol}$$

11.84
$$\ln\left(\frac{P_1}{P_2}\right) = \frac{\Delta H_{vap}}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

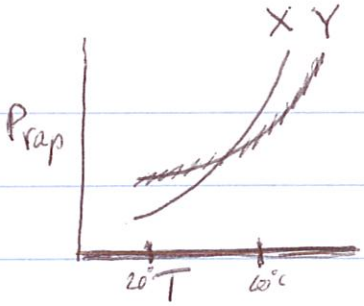
$$\ln\left(\frac{P_1}{40.1\ mmHg}\right) = \frac{31000\ J/mol}{8.31\ J/mol K}\left(\frac{1}{280.6} - \frac{1}{333.6}\right)$$

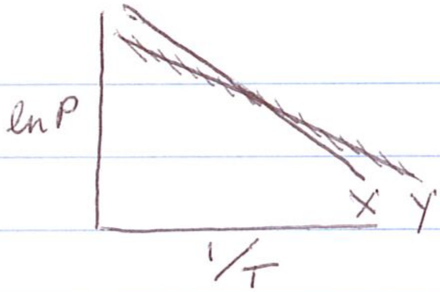
$$\ln\left(\frac{P}{40.1\ torr}\right) = 2.11214063$$

$$\frac{P}{240.1\ torr} = \ln^{-1}(2.11214063)\ \left(\text{or } e^{2.11214063}\right) = 8.266\ldots$$

$$P = (8.265916629) \cdot (40.1\ torr) = \boxed{331\ torr}$$

11.85





Since the ~~slope of~~ $\ln P_X$ v. $1/T$ has a steeper slope than $\ln P_Y$ v. $\frac{1}{T}$, X must have a greater ΔH_{vap} than Y does.

11.86 Make no mistake - both would be painful, but the liquid N_2 will take some energy as it evaporates, but it will also leave your skin as it evaporates. The boiling water will stay on your skin as it cools, transferring a large amount of energy to your skin the entire time. Also, water has a very large specific heat capacity, so it will be able to transfer a lot of energy for each degree that it cools down.