

Thermodynamics. Problems. Sheet 5.

1. Recall that material equilibrium between two systems exists when there is zero material transfer between them. Are the following composite systems in material equilibrium with their surroundings?

- | | |
|--|---|
| a) Water in a glass at ambient temperature | Not typically, unless the ambient relative humidity were very high (100% in the case of pure water). |
| b) Water in a closed bottle. | Yes, given enough time after the bottle was closed. |
| c) Water in an open bottle. | Not typically. Same case as water in a glass. |
| d) Water in our bodies. | Not, since our bodies are constantly releasing water into the ambient (through breathing and perspiration). |
| e) Hot coffee in a high-quality thermos. | Yes. An equilibrium would be reached over a short observation period. Over a very long period, the temperature would slowly diminish. Throughout this process the water partial pressure in the head space above the liquid would be changing, approximately maintaining materials equilibrium. |
| f) Ice cubes in the freezer (no-frost freezer). | No. The ambient in a no-frost freezer is dry. Ice cubes will sublime and eventually disappear. (Have you ever observed that?) |
| g) Snow on a mountain. | Not typically, unless the ambient conditions are humid enough to prevent sublimation of the snow. |
| h) Sea water. | Not typically, unless the ambient conditions are humid enough to prevent evaporation of the sea water. |
| i) Water vapor in a bubble inside Antarctic ice. | Yes. |

j) Carbon dioxide in a cola beverage before and after opening it.

Yes, before. Not, after. Notice that the closed beverage is pressurized and, when it is opened, fizzes as it releases carbon dioxide into the ambient.

2. Compute the heat of vaporization of 1g of ethylamine from the following data.

T (°C)	P (mmHg)
-22.9	111.2
-13.9	183
-5.6	281.8
5.8	481.3
16.2	750.5

[Answer: 652 J]

These data describe the phase equilibria coexistence curve in a PT diagram between the liquid and vapor phases of **ethylamine**. Therefore, the Clapeyron equation would be applicable:

$$\frac{dP}{dT} = \frac{\Delta H_v}{T \Delta v}$$

However, we do not have v_l . However, we can assume that

$$\Delta v \equiv v_g - v_l \approx v_g \approx \frac{RT}{P}$$

Under these conditions, we can apply the integrated form of the above equation, that is, the Clausius-Clapeyron equation:

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

This equation is satisfied by every pair of points, and therefore, for any pair of points, we could compute the value of ΔH_v . This, however, is not a very rigorous method, as can be seen by the results, easily obtained in this way, which are listed next:

	A	B	C	D	E	F
1	T (°C)	P (mmHg)	T(K)	P(Pa)	DH _v (J/mol)	DH _v (J/g)
2	-22.9	111.2	250.3	14825		
3	-13.9	183	259.3	24398	29856	662.2
4	-5.6	281.8	267.6	37570	29995	665.3
5	5.8	481.3	279.0	64168	29136	646.2
6	16.2	750.5	289.4	100058	28665	635.8
7						
8		R	8.314	J/mol·K		652.4
9		MW	45.085	g·mol ⁻¹		average

It is generally better to carry out a regression fit. To do so, we can find a way to reduce the Clausius-Clapeyron equation to a straight line with zero intercept:

$$Y = M \cdot x$$

where we have defined:

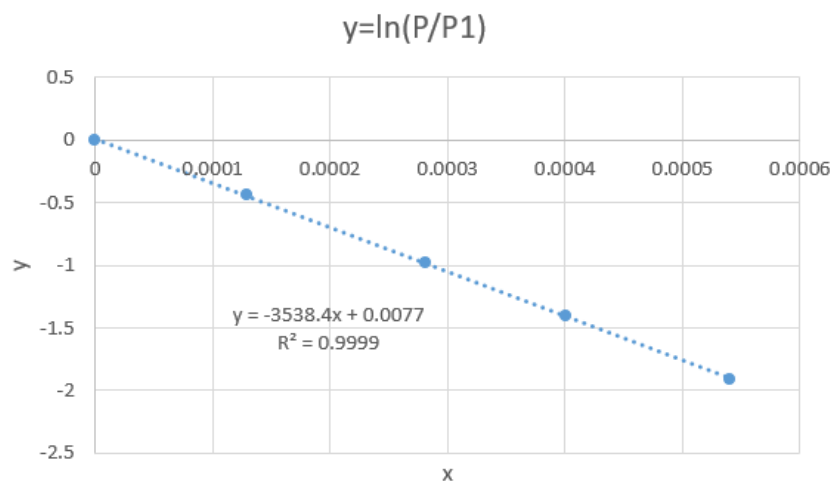
$$y \equiv \ln\left(\frac{P}{P_1}\right)$$

$$x \equiv \left(\frac{1}{T} - \frac{1}{T_1}\right)$$

$$M \equiv -\frac{\Delta H_v}{R}$$

The choice of point (P_1, T_1) is inconsequential for finding the slope, as all this choice determines is the location of the origin.

Here is the regression fit done by Excel:



Note that we did not force the ordinate at the origin ($y(x = 0)$) to be 0, as the data may contain error. From the slope we can easily compute (for 1 g) $\Delta H_v = -MR = 652.5$ J/g. The regression fit slope leads to approximately the same as the average of the piecemeal values obtained earlier. This is because the values lie very close to the fitting line.

3. The vapor pressure of 1 mole of a certain ketone is given in the following table:

T (°C)	P (Torr)
57.4	1.00
100.4	10.0
133.0	40.0
157.3	100
203.5	400
227.5	760

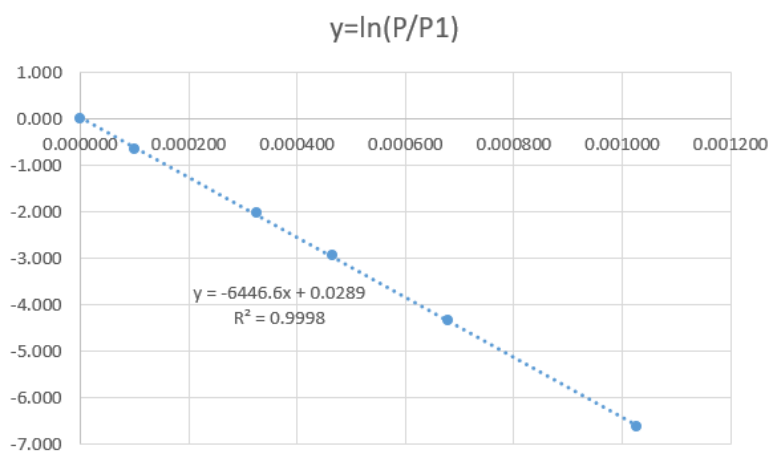
- a) Which is the normal boiling point?
- b) Which is the enthalpy of vaporization of the ketone?

Use degrees Kelvin. 1 Torr = 133.322 Pa.

[Answer: a) 227.5 °C; b) 53.6 kJ/mol]

a) The normal boiling point refers to the liquid-vapor equilibrium temperature at 1 atm pressure. We can simply get the result from the table, since 1 atm is 760 Torr, and the equilibrium temperature (boiling points) is given for this pressure.

b) To get the enthalpy of vaporization (ΔH_v), we apply the Clausius-Clapeyron equation, and carry out a regression as explained in problem 1:



From the slope we can easily compute $\Delta H_v = 53.6$ kJ/mol.

4. In a temperature range close to the transition from orthorhombic sulfur to monocline sulfur, $S_{8_orto}(s) = S_{8_monocl}(s)$, vapor pressures can be expressed by the following equations:

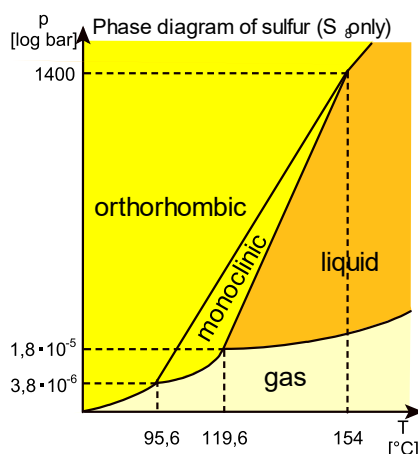
$$\log P_{ortho} = -5267 T^{-1} + 11.866$$

$$\log P_{monocl} = -5082 T^{-1} + 11.364$$

(T given in K; vapor pressures in mmHg)

Compute the temperature and the enthalpy at the transition.

[Answer: 368.5 K; 3.54 kJ/mol]



From https://commons.wikimedia.org/wiki/File:Sulfur_phase_diagram.svg

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At the transition (coexistence point of vapor and the orthorhombic and monoclinic phases), $P_{ortho} = P_{monocl}$. Therefore, we can equate and solve for T :

$$-5267/T + 11.866 = -5082/T + 11.364$$

Ergo, $T = 368.5 \text{ K (95.4}^\circ\text{C)}$. (We will not concern ourselves with the slight difference in the temperature value between the one given here and the one in the diagram.) Note we can also compute the pressure at the transition, easily by substituting this temperature in any of the equilibrium curves, and solving for $P = 3.75 \text{ Torr} = 0.50 \text{ Pa}$.

To compute the enthalpy of the orthorhombic / monoclinic transition ΔH_{ss} , we use Hess's law (see notes of topic 2). Since we are computing the heat involved in converting one phase to the other, any path which accomplishes the transition will lead to the same net heat transferred from a heat source. Therefore, we imagine a path where we go from the orthorhombic phase to the vapor phase, and then from the vapor phase to the monoclinic phase. The path is designed around the transition point and can be made as small as needed so that only the phase changes contribute to any significant heat interactions with the source. In other words, we will compute ΔH_{ss} as

$$\Delta H_{ss} = \Delta H_v^{ortho} - \Delta H_v^{monocl}$$

Where ΔH_v^{ortho} and ΔH_v^{monocl} are the heats of vaporization of the two solid phases in question, and for which we can identify the terms of the Clausius Clapeyron equation in the expressions of the problem statement, such that:

$$\frac{\Delta H_v^{ortho}}{R} = 5267 \text{ K} \times \ln(10)$$

$$\frac{\Delta H_v^{monocl}}{R} = 5082 \text{ K} \times \ln(10)$$

Therefore,

$$\Delta H_{ss} = \Delta H_v^{ortho} - \Delta H_v^{monocl} = (5267 - 5082)(\text{K}) \times R \times \ln(10) = \mathbf{3.54 \text{ kJ/mol}}$$

where we've used that $\log P = \ln(P) / \ln(10)$.

5. At 0°C and 1.00 atm, the molar volume of water is 18.018 cm³ and that of ice is 19.64 cm³. How would the melting point shift if the pressure was increased to 3.00 atm? The melting heat of ice is 79.69 cal/g.

[Answer: -0.015 K]

We are asked to compute the temperature of melting T at a pressure of 3 atm. We can apply the Clapeyron equation:

$$\frac{dP}{dT} = \frac{\Delta H_m}{T \Delta v_{ls}}$$

Note that we cannot apply the Clapeyron equation by taking the differentials as finite differences, since the pressure and temperature relationship is not linear. We need to integrate this equation between the two states considered: (273.5 K, 101325 Pa) and (T , 3×101325 Pa). We write in the following form,

$$dP = \frac{\Delta H_m}{\Delta v_{ls}} \frac{dT}{T}$$

And integrate assuming that the heat of melting and molar volume difference stay constant:

$$\int_{P_1=1 \text{ atm}}^{P_2=3 \text{ atm}} dP = \frac{\Delta H_m}{\Delta v_{ls}} \int_{T_1=273.15 \text{ K}}^{T_2} \frac{dT}{T}$$

Ergo,

$$\ln\left(\frac{T_2}{T_1}\right) = \frac{\Delta v_{ls}}{\Delta H_m} (P_2 - P_1)$$

$$T_2 = T_1 e^{\frac{\Delta v_{sl}}{\Delta H_m} (P_2 - P_1)}$$

$$T_2(\text{K}) = 273.15 \times e^{\frac{-1.622 \times 10^{-6}}{6.006 \times 10^3} \times (3-1) \times 101325} =$$

$$T_2 = 273.135 \text{ K}$$

Where we have used international units, and have used that

$$\Delta v_{ls} = v_l - v_s = (18.018 - 19.64) \times 10^{-6} \frac{\text{m}^3}{\text{mol}} = -1.622 \times 10^{-6} \frac{\text{m}^3}{\text{mol}}$$

and have expressed the heat of melting per mol instead of per gram, by multiplying by the molar mass of water (18.015 g/mol):

$$\Delta H_m = 79.69 \left(\frac{\text{cal}}{\text{g}} \right) * 4.184 \left(\frac{\text{J}}{\text{cal}} \right) * 18.015 \left(\frac{\text{g}}{\text{mol}} \right) = 6006 \frac{\text{J}}{\text{mol}}$$

Finally, the melting point shift (change in temperature) is computed to be:

$$\Delta T \equiv T_2 - T_1 = 273.135 - 273.15 = -0.015 \text{ K}$$

It is important to note that this value assumes an accuracy in the final temperature of 4 parts per million approximately ($=1/273135=3.7 \times 10^{-6}$), but that the data provided in the problem statement do not require this level of accuracy. This statement can be easily checked by changing the input values slightly in a spreadsheet.

6. Trouton's rule states that the entropy of vaporization of most liquids is approximately equal to 21.5 cal/(K·mol), at boiling temperature and pressure of 1 atm. Compute by means of an approximate expression the change in the point boiling resulting from a drop in the pressure from 760 to 750 mmHg, for toluene, whose normal boiling point is 110.6°C.

[Answer: -0.47 K]

To compute the change in the temperature for a change in the pressure we need the slope of the coexistence liquid-vapor curve for **toluene**. We can compute this from the Clausius-Clapeyron equation:

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

They are asking us for $T_2 - T_1$ (change in the boiling point; that is, boiling temperature) where $T_1 = (110.6 + 273.15) \text{ K}$, $P_2 = 750 \text{ mmHg}$, and $P_1 = 760 \text{ mmHg}$. Therefore T_2 can be expressed as:

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

To get the heat (enthalpy) of vaporization, we use the entropy of vaporization. The two are connected by the following expression, which is the integral form of $dH = TdS - VdP$ where we consider the phase transition to take place at constant pressure, thus $-VdP = 0$, and constant temperature T :

$$\Delta H = T\Delta S$$

(We are considering molar quantities.) Therefore,

$$\Delta H = T\Delta S = (110.6 + 273.15)\text{K} \times 21.5 \frac{\text{cal}}{\text{K} \cdot \text{mol}} \times 4.184 \frac{\text{J}}{\text{cal}} = 34.521 \frac{\text{kJ}}{\text{mol}}$$

This value is assumed to be the same at the two states (1 and 2), as they are very near to each other ($T_2 \approx T_1$). There is no point in getting a better value by (for example) iterating the solution, since the entropy in the Trouton's rule is an approximate number.

Substituting in the earlier expression (after noting that no conversion for the pressure units is needed):

$$\frac{1}{T_2} = \frac{1}{110.6 + 273.15} - \frac{8.314}{34521} \ln \left(\frac{760750}{750760} \right) = 2.609 \times 10^{-3} \text{ K}^{-1}$$

$$T_2 = 383.28 \text{ K}$$

$$T_2 - T_1 = 383.28 - 383.75 = -0.47 \text{ K}$$

Another way to solve this problem:

Note that since the pressure change is very small, we would expect a small temperature change, to be checked a posteriori. Therefore, we can approximate the change in temperature by taking the differentials in the Clapeyron equation as finite increments, from

$$\frac{dP}{dT} = \frac{\Delta H_v}{T \Delta v}$$

to

$$\frac{\Delta P}{\Delta T} \approx \frac{\Delta H_v}{T \Delta v}$$

From here, and after replacing Δv with RT/P (which neglects the molar volume of the liquid phase, and assumes gas-ideality for the vapor),

$$\Delta T \approx \frac{RT_1^2 \Delta P}{P_1 \Delta H_v} = -0.47 \text{ K}$$

We indeed get a small decrement in temperature, as we had assumed. Therefore, the result must be correct. Indeed, we got the same result as by the more exact procedure (at least if expressed by two significant figures).

Additional comment: Note the remarkable fact condensed in Trouton's rule. As physicists we may ask why would the entropy involved in vaporization be nearly the same per mole (or number of molecules) independently of the substance (e.g., its molecular weight). Why should it be per mole and not per mass or some other basis?

7. Compute the pressure required to melt 1.00 g of ice at a temperature of -1°C , knowing that its heat of melting is 79.69 cal/g . The density of ice is 0.92 g/cm^3 and that of liquid water is 1.00 g/cm^3 .

[Answer: 139 atm]

To compute the required pressure, we consider a known (P, T) point on the ice-water coexistence curve, for example $(P_1=1 \text{ atm}, T_1=0^{\circ}\text{C})$, and then we extrapolate to the new temperature assuming that the coexistence curve has a constant slope between the points. This is a good approximation as the heat of melting and the densities almost do not change between $T_1 = 273.15 \text{ K}$ and $T_2 = -1^{\circ}\text{C} = 272.15$, and so are given as having fixed values in the problem statement. We consider the Clapeyron equation:

$$\frac{dP}{dT} = \frac{\Delta H_m}{T \Delta v_{ls}}$$

where the enthalpy and volume differences are referred to 1 g:

$$\Delta H_m = 79.69 \frac{\text{cal}}{\text{g}} \times 4.184 \frac{\text{J}}{\text{cal}} = 333.42 \frac{\text{J}}{\text{g}}$$

$$\Delta v_{ls} \equiv v_l - v_s = \left[\left(\frac{1}{1.00} \right) - \left(\frac{1}{0.92} \right) \right] \frac{\text{cm}^3}{\text{g}} = -8.696 \times 10^{-8} \frac{\text{m}^3}{\text{g}}$$

Therefore, since the difference in temperature is small, we can avoid doing the integral, and can estimate the slope as:

$$\frac{dP}{dT} = \frac{\Delta H_m}{T \Delta v_{ls}} \approx \frac{\Delta H_m}{T_1 \Delta v_{ls}}$$

Finally,

$$\frac{P_2 - P_1}{T_2 - T_1} \approx \frac{\Delta H_m}{T_1 \Delta v_{ls}}$$

$$P_2 = P_1 + (T_2 - T_1) \frac{\Delta H_m}{T_1 \Delta v_{ls}}$$

$$P_2 = 101325 + (-1) \frac{333.42}{273.15 \times (-8.696) \times 10^{-8}} = \mathbf{1.414 \times 10^7 \text{ Pa} = 139 \text{ atm}}$$

Note that the result is quite sensitive on the accuracy of the densities. If we had used the more “accurate” values of 0.9998 and 0.917 g/cm^3 for the densities of water and ice, respectively, we would have gotten a value of P_2 of 134.4 atm , closer to the real value.

8. The vapor pressure of benzene at 21°C is 76 Torr and its normal temperature of boiling is 80°C. Compute the enthalpy of vaporization explaining the approximations used.

[Answer: 33.7 kJ/mol]

The two data points are points on the phase equilibria coexistence curve in a $P - T$ diagram between the liquid and vapor phases of benzene. Therefore, the Clapeyron equation would be applicable:

$$\frac{dP}{dT} = \frac{\Delta H_v}{T \Delta v}$$

Although we do not have v_l it is much smaller than ... so we can neglect it, such that

$$\Delta v \equiv v_g - v_l \approx v_g$$

In addition, we assume that the vapor phase behaves as an ideal gas over the temperature and pressure ranges considered,

$$v_g \approx \frac{RT}{P}$$

With these assumptions, we can integrate the Clapeyron equation, to obtain the Clausius-Clapeyron equation:

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

where we have assumed that ΔH_v stays constant between the two states considered. From here we can compute ΔH_v as:

$$\Delta H_v = R \left(\frac{T_2 T_1}{T_2 - T_1} \right) \ln\left(\frac{P_2}{P_1}\right) = 33.7 \frac{\text{kJ}}{\text{mol}}$$

This is quite close to the heat of vaporization of benzene at normal conditions of 33.9 kJ/mol. See: <https://webbook.nist.gov/cgi/inchi?ID=C71432&Mask=4>

9. Over a wide temperature range, the vapor pressures (in mmHg) of solid and liquid UF_6 follow the approximate equations:

$$\log P_s = 10.648 - 2559.5 T^{-1}, \text{ and}$$

$$\log P_L = 7.5396 - 1511.3 T^{-1}.$$

- At what condition of temperature and pressure can solid, liquid and gas UF_6 coexist in equilibrium?
- At which temperature is UF_6 in equilibrium with its vapor at 1 atm? At this condition, is the condensed phase solid or liquid?

[Answer: a) 337 K; 1.50 atm; b) 329 K; solid]

a) At the triple point solid, liquid and vapor coexist, therefore both equations will be fulfilled. Therefore, we can equate the left-hand sides and solve for $T (= T_3)$:

$$-2559.5/T + 10.648 = -1511.3/T + 7.5396$$

Ergo, $T_3 = 337.2 \text{ K} (64^\circ\text{C})$. Then, by substituting this temperature in either of the vapor pressure curves, and solving for P we get $P_3 = 1142 \text{ mmHg} = 1.50 \text{ atm}$.

b) Since $P = 1 \text{ atm} (760 \text{ mmHg})$ is below the triple point pressure of 1.50 atm, the condensed phase in equilibrium with the vapor will be the solid phase. Therefore, we can obtain the temperature of equilibrium by substituting 760 mmHg in the equation for the solid-vapor equilibrium $P_s(T)$

$$\log(760) = -2559.5/T + 10.648$$

and then solve for T , to obtain: $T = 329.5 \text{ K} = 56.4^\circ\text{C}$.

10. In a temperature range close to the triple point of TaBr_5 , the vapor pressures (in mmHg) of the solid and the liquid follow the following equations:

$$\log P_s = 12,571 - 5650 T^{-1}, \text{ and}$$

$$\log P_L = 8,171 - 3265 T^{-1}.$$

Compute for TaBr_5

- the triple point and ΔH of melting, and
- the boiling point and the entropy of evaporation.

[Answer: a) 542 K; $4.57 \times 10^4 \text{ J/mol}$; b) 617 K; 101.3 J/K mol]

a) The first question is easy: The **triple point** is the point where the three coexistence curves meet. Therefore, in particular, the two vapor-condensed phase curves will be satisfied simultaneously. Thus, we can equate the right-hand sides of the expressions in the problem statement, and solve for T_3 , and then substitute the value in any of the two equations to find P_3 , as follows:

$$12.571 - \frac{5650}{T_3} = 8.171 - \frac{3265}{T_3}$$

$$\log(P_3) = 12.571 - \frac{5650}{T_3}$$

Therefore, the triple point is at:

$$T_3 = 542.04 \text{ K}$$

$$P_3 = 140.45 \text{ mmHg}$$

The enthalpy of melting (AKA heat of melting) ΔH_m is related to the slope of the P vs. T solid-liquid coexistence curve, which is given by the Clapeyron equation:

$$\frac{dP}{dT} = \frac{\Delta H_m}{T \Delta v_{ls}}$$

However, the necessary unknowns to solve for ΔH_m are not available ($\frac{dP}{dT}$ and Δv_{ls}).

Instead, we have the dependences for the other two coexistence curves (solid/vapor and liquid/vapor).

To solve the problem, we need to apply Hess's law (see notes of topic 2) to a path that takes us from the solid phase to the liquid phase ($s \rightarrow l$) in two steps: from solid to gas ($s \rightarrow g$), and from gas to liquid ($g \rightarrow l$). The net enthalpy change (per 1 mol) will then be the heat of melting:

$$\Delta H_m \equiv \Delta H_{ls} = H_l - H_s = (H_g - H_s) + (H_l - H_g)$$

$$\Delta H_{ls} = \Delta H_{gs} + \Delta H_{lg} = \Delta H_{gs} - \Delta H_{gl}$$

Where, we have introduced ΔH_{ls} to rename ΔH_m to help understand the enthalpy changes along the path.

We can now express the enthalpy terms on the right-hand side of the above expression as a function of known parameters, by means of the Clapeyron equation applied to the $s \rightarrow l$ and $l \rightarrow g$ phase transitions,

$$\frac{dP_s}{dT} = \frac{\Delta H_{gs}}{T \Delta v_{gs}}$$

$$\frac{dP_l}{dT} = \frac{\Delta H_{gl}}{T \Delta v_{gl}}$$

But first we must introduce the approximation that the change in molar volumes is approximately equal to the molar volume of the gas phase $\Delta v_{gx} \approx v_g$ ($x = s$ or l), and that such volume is correctly expressed by the ideal gas law $v_g = RT/P$:

$$\frac{dP}{dT} = \frac{P \Delta H_{gx}}{R T^2}$$

For $x = s$ or l . Now we transform this expression, as

$$\frac{d \ln P}{dT} = \frac{\Delta H_{gx}}{R T^2}$$

Since

$$\log(P) = \ln(P) / \ln(10)$$

$$\frac{d \log P}{dT} = \frac{\Delta H_{gx}}{\ln(10) R T^2}$$

We can now identify the coefficients to the T^{-1} terms in the problem statement equations with

$$-\frac{\Delta H_{gx}}{\ln(10) R}$$

Note that the units of these terms are kelvin. Finally, we have that:

$$\Delta H_{gs} = \ln(10) \cdot R \cdot 5650 = 108.16 \frac{\text{kJ}}{\text{mol}}$$

$$\Delta H_{gl} = \ln(10) \cdot R \cdot 3265 = 62.50 \frac{\text{kJ}}{\text{mol}}$$

Therefore,

$$\Delta H_m \equiv \Delta H_{ls} = \Delta H_{gs} - \Delta H_{gl} = 108.16 - 62.50 = \mathbf{45.65 \frac{kJ}{mol}}$$

b) Let's now get the boiling point and the entropy of vaporization, which we need to interpret at 1 atm. To find the boiling point, we need to find the temperature along the liquid-vapor coexistence curve, where $P = 1 \text{ atm} = 760 \text{ mmHg}$. Therefore, we must solve the following expression for T:

$$\log(760) = 8.171 - \frac{3265}{T_b}$$

Ergo,

$$T_b = 617 \text{ K}$$

To find the entropy of vaporization, at the boiling point, we need simply to compute:

$$\Delta S_{gl} = \frac{\Delta H_{gl}}{T_b} = \frac{62.50 \text{ kJ}}{617 \text{ K} \cdot \text{mol}} = \mathbf{101.3 \frac{J}{K \cdot \text{mol}}}$$

11. The melting heat of naphthalene at the normal melting point of 80.0°C is 36.0 cal/g. Compute the change in the melting point when the pressure increases by 1 atm, knowing that the density of solid naphthalene is 1.145 g/cm³ and that of liquid naphthalene is 0.981 g/cm³.

[Answer: +0.035 K]

The initial and the final states are points on the solid-liquid coexistence curve, which satisfies the Clapeyron equation:

$$\frac{dP_{ls}}{dT} = \frac{\Delta H_{ls}}{T (v_l - v_s)}$$

where, ΔH_{ls} , v_l and v_s are assumed constant and are given in the problem:

$$\Delta H_{ls} = 36.0 \frac{\text{cal}}{\text{g}} = 36.0 \frac{\text{cal}}{\text{g}} \cdot 4.184 \frac{\text{J}}{\text{cal}} = 150.6 \frac{\text{J}}{\text{g}}$$

$$v_l - v_s = \frac{1}{0.981} - \frac{1}{1.145} = 0.1460 \frac{\text{cm}^3}{\text{g}} = 1.460 \times 10^{-7} \frac{\text{m}^3}{\text{g}}$$

They are expressed per the same unit mass (g); where the choice of this unit taken does not matter as it will cancel out.

If these are constant, we can integrate the Clapeyron equation along the curve between two states 1 and 2, as shown next:

$$\int_{P_1}^{P_2} dP_{ls} = \frac{\Delta H_{ls}}{(v_l - v_s)} \int_{T_1}^{T_2} \frac{dT}{T}$$

$$P_2 - P_1 = \frac{\Delta H_{ls}}{(v_l - v_s)} \ln\left(\frac{T_2}{T_1}\right)$$

The unknown of the problem is the final temperature, T_2 . Therefore, we can solve the above expression for T_2 as:

$$T_2 = T_1 e^{\frac{(P_2 - P_1)(v_l - v_s)}{\Delta H_{ls}}}$$

The initial and final (2) pressures P_1 and P_2 and initial temperature T_1 are given by:

$$P_1 = 1 \text{ atm} = 101325 \text{ Pa}$$

$$P_2 = P_1 + 1 \text{ atm} = 2 \times 101325 \text{ Pa}$$

$$T_1 = 80^\circ\text{C} = 353.15 \text{ K}$$

After substituting the values in the above expression for T_2 , we get:

$$T_2 (\text{K}) = 353.15 \times e^{\frac{(2-1) \times 101325 \times (1.460 \times 10^{-7})}{150.6}} = 353.1847 \text{ K}$$

The change in the melting point is

$$\Delta T \equiv T_2 - T_1 = 353.1847 - 353.15 = +0.035 \text{ K}$$