Recitation 2

For the following problems, you may need physical constants that aren't listed. They can be found either in your book or on the internet, or ask and we'll find them together.

Piece #1 What is the mass (in g) of 250.0 mL of ethanol (C_2H_5OH) at room temp (298 K)?

On page 18, density of ethanol is 0.789 g/cm^3 at 293 K.

$$250.0\ mL\ C_2H_5OH \frac{1\ cm^3}{1\ mL} \frac{0.789\ g}{cm^3 C_2H_5OH} = 197.25\ g\ C_2H_5OH$$

Technically, 197 g to the right number of significant figures.

Piece #2 How many moles of ethanol are in 250.0 mL at room temp (298 K)?

$$250.0\ mL\ C_2H_5OH \frac{1\ cm^3}{1\ mL} \frac{0.789\ g}{cm^3 C_2H_5OH} = 197.25\ g\ C_2H_5OH \frac{1\ mol\ C_2H_5OH}{46.07\ g} = 4.28\ mol\ C_2H_5OH$$

Piece #3 How much energy (in Joules) is required to heat 250.0 mL of ethanol from 298 K to 310 K?

We need to be a little careful about phase changes.

The melting point of ethanol is -114.1°C (159.05 K)

The boiling point of ethanol is 78.3°C (351.45 K)

So, the ethanol is a liquid the whole time and I'm just heating it up.

$$Q = mc\Delta T = (197.25\ g)\left(2.42 \frac{J}{gK}\right)(310\ K - 298\ K) = 5728.14\ J$$

Piece #4 How much energy (in Joules) is required to boil the ethanol from Piece #3 if it is already at its boiling point?

Phase changes are all about the enthalpy change.

From Wikipedia:

ΔH_{fus} = 4.9 kJ/mol

ΔH_{vap} = 38.56 kJ/mol

$$Q = n\Delta H_{vap} = (4.28\ mol)\left(38{,}560 \frac{J}{mol}\right) = 1.65 \times 10^5 J$$

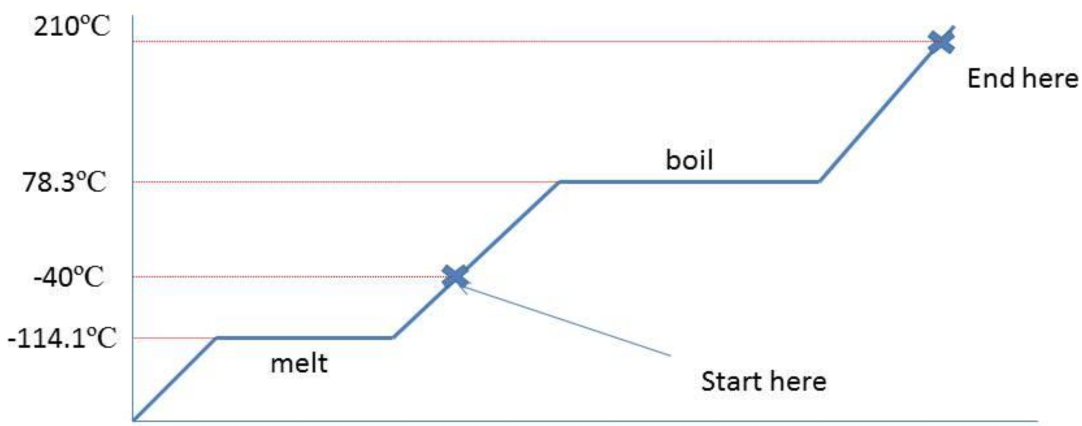
Puzzle #2 How much energy would it take to heat 250.0 mL of ethanol (C_2H_5OH) from -40ºC to 210ºC?

You need to consider what happens along the path from -40ºC to 210ºC. Essentially, we are looking at the phase diagram at 1 atm (standard pressure).

The normal melting point of ethanol is -114.1°C (159.05 K)

The normal boiling point of ethanol is 78.3°C (351.45 K)

Since we start at -40ºC, above the melting point, the ethanol is a liquid when we start. It will stay liquid until we reach 78.3°C at which point it boils. Then it is a gas from 78.3 °C to 210 °C. So my heating curve would look like:



So, we have 3 things going on:

1. We heat the liquid ethanol to its boiling point (78.3ºC)

2. We boil the ethanol.
3. We heat the gaseous ethanol to its final temperature (210 ºC)

Heating is: Q=mcΔT

Boiling is: $Q=n\Delta H_{vap}$

$$Q_{total} = Q_{heat\ liquid} + Q_{boil} + Q_{heat\ gas}$$

$$Q_{total} = m_{eth}C_{liq,eth}\Delta T + n_{eth}\Delta H_{vap,eth} + m_{eth}C_{gas,eth}\Delta T$$

We have the numbers from the “pieces” of the puzzle earlier.

$$Q_{total} = (197.25\ g)\left(2.42\frac{J}{gK}\right)(78.3℃ - (-40℃)) + (4.28\ mol\ eth)\left(38{,}560\frac{J}{mol}\right) + (197.25\ g)\left(1.7\frac{J}{gK}\right)(210℃ - 78.3℃) = 265{,}669\ J$$

Piece #1 What is the Normal Boiling Point of ethanol (in kelvins)?

78.3 ºC

Piece #2 What is the enthalpy of vaporization of ethanol at its normal boiling point?

38,560 J/mol

Puzzle #3 What would the boiling point of ethanol be at the top of Mt. Everest where the average atmospheric pressure is 0.64 atm?

Clausius Clapeyron:

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

Lots of variables, but we know most of them:

We know the normal boiling point which is the boiling point at 1 atm, so:

P_1=1 atm

T_1=78.3ºC

ΔH_{vap} = 38,560 J/mol

R=8.314 J/mol K

P_2 = 0.64 atm at top of Everest

T_2 = ? Boiling point at 0.64 atm.

T_2 is the only thing we don't know.

As always, be careful of UNITS! UNITS! UNITS!

T must be in Kelvins

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

$$\ln\left(\frac{1\ atm}{0.64\ atm}\right) = \frac{-38{,}560\frac{J}{mol}}{8.314\frac{J}{molK}}\left(\frac{1}{(78.3 + 273.15\ K} - \frac{1}{T_2}\right)$$

We just need to solve for T_2:

$$0.4463 = -4638\left(\frac{1}{351.45} - \frac{1}{T_2}\right)$$

$$-9.622 \times 10^{-5} = \left(0.002845 - \frac{1}{T_2}\right)$$

$$-0.002941 = -\frac{1}{T_2}$$

$$T_2 = 340$$

Which makes sense. 340<351. So, when the pressure went down, so did the boiling point.

Puzzle #4 I want ethanol to boil at room temperature (298 K). What would I need to reduce the pressure to in order to accomplish this?

Clausius Clapeyron again!

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

P_1=1 atm

T_1=78.3ºC

ΔH_{vap} = 38,560 J/mol

R=8.314 J/mol K

P_2 = ? I don't know

T_2 = 298 K (my desired boiling point)

$$\ln\left(\frac{1\ atm}{x\ atm}\right) = \frac{-38{,}560\,\frac{J}{mol}}{8.314\,\frac{J}{molK}}\left(\frac{1}{(78.3 + 273.15\ K} - \frac{1}{298\ K}\right)$$

$$\ln\left(\frac{1}{x}\right) = -4638\ K^{-1}(0.002845 - 0.003356) = 2.367$$

$$\ln\left(\frac{1}{x}\right) = 2.367$$

$$e^{\ln\left(\frac{1}{x}\right)} = e^{2.367}$$

$$\frac{1}{x} = 10.67$$

$$x = 0.094\ atm$$