3.2) Calculate the heat required to evaporate 1.00 kg of water at (a) 25°C, (b) 100°C.

The process is: \( \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(g) \)

\[
1.00 \times 10^3 \text{ g} \times \frac{1 \text{ mol}}{18.016 \text{ g}} = 55.506 \text{ mol}
\]

From Table 3.1, \( \Delta H_{\text{vap}}^0 (373.15 \text{ K}) = 40.7 \text{ kJ/mol} \)

For 298.15 K, use heat of formation data at 298.15 K to obtain the enthalpy of the above reaction, which is equal to the enthalpy of vaporization:

\[
\Delta H = \Delta H_f^0 (\text{H}_2\text{O}(g)) - \Delta H_f^0 (\text{H}_2\text{O}(l)) = 44.01 \text{ kJ/mol}
\]

(a) at 25°C, \( q = \Delta H = (55.506)(44.01) = 2.44 \times 10^3 \text{ kJ} \)

(b) at 100°C, \( q = (55.506)(40.7) = 2.26 \times 10^3 \text{ kJ} \)

3.5) Use the information in Tables 2.1 and 3.1 to calculate the total heat required to melt 100 g of ice at 0°C, heat it to 100°C, and then vaporize it at that temperature. Sketch a graph of temperature against time on the assumption that heat is supplied to the sample at a constant rate.

From the above tables:

\( C_{p,m}(\text{H}_2\text{O}(l)) = 75.29 \text{ kJ/mol} \)

\( \Delta H_{\text{fus}}(\text{H}_2\text{O}) = 6.01 \text{ kJ/mol at 0°C} \)

\( \Delta H_{\text{vap}}(\text{H}_2\text{O}) = 40.7 \text{ kJ/mol at 100°C} \)

\[
n = 100 \text{ g} \text{H}_2\text{O}(s) \times \frac{1 \text{ mol}}{18.016 \text{ g}} = 5.5506 \text{ mol H}_2\text{O}(s)
\]

The process we’re interested in is:

\( \text{H}_2\text{O}(s), 0°C \rightarrow \text{H}_2\text{O}(l), 0°C \rightarrow \text{H}_2\text{O}(l), 100°C \rightarrow \text{H}_2\text{O}(g), 100°C \)
The total heat required is the sum of 3 enthalpy changes:

\[ q = n\Delta H^0 = n\Delta H_{fus}^0 + nC_p,m \Delta T + n\Delta H_{vap}^0 \]

\[ = (5.5506)[6.01 + (0.07529)(100) + 40.7] = 301 \text{ kJ} \]

Assuming constant rate of heating (approximately to scale):

3.6) The enthalpy of sublimation of calcium at 25°C is 178.2 kJ/mol. How much energy (at constant temperature and pressure) must be supplied to 10.0 g of solid calcium to produce a gas composed of Ca\(^{2+}\) ions and electrons?\)

The overall process of interest is: \( \text{Ca}(s) \rightarrow \text{Ca}^{2+}(g) + 2e^- \)

But the path we will take is: \( \text{Ca}(s) \rightarrow \text{Ca}(g) \rightarrow \text{Ca}^+(g) + e^- \rightarrow \text{Ca}^{2+}(g) + 2e^- \)

The 1st step is just the enthalpy of sublimation, while the next two are \( \Delta H_{\text{ion,1}} = 590 \) kJ/mol and \( \Delta H_{\text{ion,2}} = 1150 \) kJ/mol (Table 3.2)

\[ n = 10.0 \text{ g Ca(s)} \times \frac{1 \text{ mol}}{48.08 \text{ g}} = 0.249501 \text{ mol Ca(s)} \]
\[ \Delta H = (0.249501)(178.2 + 590 + 1150) = 479 \text{ kJ} \]

3.10) The standard enthalpy of reaction for \(2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)\) is –3120 kJ/mol. (a) What is the standard enthalpy of combustion of ethane?
(b) What is the change in enthalpy when 3.00 mol of CO\(_2\) is formed in the reaction?
(a) The reaction as written above is identical to the combustion reaction, except that it is written for 2 moles of ethane rather than for one.
Hence, \(\Delta H^\circ_c = \frac{\Delta H^\circ_r}{2} = -\frac{3120}{2} = -1560 \text{ kJ/mol} \)
(b) for 3 moles of CO\(_2\) produced, \(\Delta H^\circ_r = \frac{3}{4} \Delta H^\circ_r = -2340 \text{ kJ} \)

3.19) The standard enthalpy of decomposition of the yellow complex NH\(_3\)SO\(_2\) into NH\(_3\) and SO\(_2\) is +40 kJ/mol. Calculate the standard enthalpy of formation of NH\(_3\)SO\(_2\).
The decomposition reaction refers to: \(\text{NH}_3\text{SO}_2 \rightarrow \text{NH}_3(g) + \text{SO}_2(g)\)
Using enthalpies of formation:
\[ \Delta H^\circ_{\text{decomp}} = \Delta H^\circ_f (\text{NH}_3(g)) + \Delta H^\circ_f (\text{SO}_2(g)) - \Delta H^\circ_f (\text{NH}_3\text{SO}_2(g)) \]
or,
\[ \Delta H^\circ_f (\text{NH}_3\text{SO}_2(g)) = \Delta H^\circ_f (\text{NH}_3(g)) + \Delta H^\circ_f (\text{SO}_2(g)) - \Delta H^\circ_{\text{decomp}} \]
\[ = (-46.11) + (-296.83) - 40 \]
\[ = -383 \text{ kJ/mol} \]
3.22) The mass of a typical sugar (sucrose) cube is 1.5 g. Calculate the energy released as heat when a cube is burned in air. To what height could a person of mass 68 kg climb on the energy a cube provides assuming 20% of the energy is available for work?

\[ n = \frac{1.5 \text{ g sucrose} \times 1 \text{ mol}}{342.30 \text{ g}} = 0.0043821 \text{ mol} \]

\[ q = n\Delta H^o_c = (0.0043821)(-5645) = -24.74 \text{ kJ released} \]

20% of this: \((0.20)(-24.74) = 4.948 \text{ kJ available for work} \)

\[ w = mgh \quad \text{or} \quad h = \frac{4.948 \times 10^3 \text{ J}}{(68 \text{ kg})(9.81 \text{ m/s})} = 7.4 \text{ m} \quad \text{(better eat 2 or 3 ….)} \]

3.23) The standard enthalpy of combustion of propane gas is \(-2220 \text{ kJ/mol}\) and the standard enthalpy of vaporization of the liquid is \(+15 \text{ kJ/mol}\). Calculate (a) the standard enthalpy and (b) the standard internal energy of combustion of the liquid.

The process of interest is: \(\text{C}_3\text{H}_8(l) + 5\text{O}_2(g) \rightarrow 3\text{CO}_2(g) + 4\text{H}_2\text{O}(l)\)

(a) Use the path: \(\text{C}_3\text{H}_8(l) + 5\text{O}_2(g) \rightarrow \text{C}_3\text{H}_8(g) + 5\text{O}_2(g) \rightarrow 3\text{CO}_2(g) + 4\text{H}_2\text{O}(l)\)

The first step is just the heat of vaporization of the liquid and the 2nd is the heat of combustion of the gas:

\[ \Delta H^o_c(\text{C}_3\text{H}_8(l)) = 15 + (-2220) = -2205 \text{ kJ / mol} \]

(b) The internal energy is related to the enthalpy by

\[ \Delta U = \Delta H - \Delta(pV) = \Delta H - \Delta(nRT) = \Delta H - RT\Delta n_g \], where \(\Delta n_g\) is the change in

the number moles of perfect gas in the process at constant T. In the combustion of the liquid, \(\Delta n_g = 3 - 5 = -2\) moles (see overall reaction above)

\[ \Delta U^o_c = (-2205) - (-2)(8.31451 \times 10^{-3})(298.15) = -2200 \text{ kJ/mol} \]
3.25) Use the data in Appendix 1 to calculate the standard enthalpies of the following reactions:

in each case use \( \Delta H_r^o = \sum_{prod} \nu_p \Delta H_f^o (prod) - \sum_{reac} \nu_r \Delta H_f^o (reac) \)

(a) \( 2\text{NO}_2(g) \rightarrow \text{N}_2\text{O}_4(g) \)

\[ (9.16 \text{ kJ/mol}) - 2(33.18 \text{ kJ/mol}) = -57.20 \text{ kJ/mol} \]

(b) \( \text{NO}_2(g) \rightarrow \frac{1}{2} \text{N}_2\text{O}_4(g) \)

this is just 1/2 that of (a), \((-57.20 \text{ kJ/mol})/2 = -28.60 \text{ kJ/mol} \)

(c) \( 3\text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow 2\text{HNO}_3(aq) + \text{NO}(g) \)

\[ (90.25) + 2(-207.36) - 3(33.18) - (-285.83) = -138.18 \text{ kJ/mol} \]

(d) Cyclopropane(g) \(\rightarrow\) propene(g)

\[ (20.42) - (53.30) = -32.88 \text{ kJ/mol} \]

(e) \( \text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l) \)

note: HCl and NaOH are strong acids and bases and are completely dissociated in water.

\[ (\text{H}_2\text{O}) + (\text{Na}^+ + \text{Cl}^-) - (\text{H}^+ + \text{Cl}^-) - (\text{Na}^+ + \text{OH}^-) \]

\[ (-285.83) + (-240.12 - 167.16) - (0 - 167.16) - (-240.12 - 229.99) = -55.84 \text{ kJ/mol} \]

3.27) Use the information in Appendix 1 to predict the standard reaction enthalpy of \( 2\text{NO}_2(g) \rightarrow \text{N}_2\text{O}_4(g) \) at 100°C from its value at 25°C.

Kirchhoff’s “Law” : \( \Delta H_r^o (T_2) = \Delta H_r^o (T_1) + \Delta_r C_p \Delta T \)

For \( \Delta H_r^o (298) \), use result from 3.25(a) : -57.20 kJ/mol

From Appendix 1,

\[ C_{p,m}(\text{N}_2\text{O}_4) = 77.28 \text{ J K}^{-1} \text{ mol}^{-1} \text{ and } C_{p,m}(\text{NO}_2) = 37.20 \text{ J K}^{-1} \text{ mol}^{-1} \]

\[ \Delta_r C_p = 77.28 - 2(37.20) = 2.88 \text{ J K}^{-1} \text{ mol}^{-1} \]

\[ \Delta H_r^o (298) = -57.20 + (2.88 \times 10^{-3})(373 - 298) = -56.98 \text{ kJ/mol} \]