Solutions to:
Energy and Enthalpy Homework Problem Set
Chemistry 145, Chapter 6

1. An electric stove has a 750 watt heating element. If it takes 5 minutes fr a teakettle to come to a boil, how much energy (in joules) is required?

$power := 750 \cdot watt$ $\quad$ $watt = 1.000 \cdot joule \cdot sec^{-1}$ $\quad$ $power = 750.000 \cdot joule \cdot sec^{-1}$

$time := 5 \cdot min$ $\quad$ $min = 60.000 \cdot sec$ $\quad$ $time = 300.000 \cdot sec$

$energy := power \cdot time$

$energy = 2.250 \cdot 10^{5} \cdot joule$

Note: Pay attention to the units here. The problem gives the power of the element in watts (a watt is equivilent to 1 joule per second) and the total time in min (a minute is 60 seconds). The energy is the power (watt) times the time (second).

2. A piece of titanium metal (mass 852.398 g) is placed in poiling water (100.0 ºC). After 20 minits it is removed from the boiling water and placed in a 1.00 liter container of water at 20.00 ºC. The temperature of the water incrreases to 27.70 ºC. What is the specific heat capacity of titanium? How does this compare to the value given in the *CRC Handbook of Chemistry and Physics*?

$mass_{metal} := 852.398 \cdot gm$

$volume_{water} := 1.00 \cdot liter$

NOTE: Mathcad uses gm as the abbreviation for gram and uses Kelvin for temperature units. This problem may also be worked using temperature in Celcius.

$T_{metal} := (273.15 + 100.0) \cdot K$ $\quad$ $T_{metal} = 373.150 \cdot K$

$T_{water} := (273.15 + 20.00) \cdot K$ $\quad$ $T_{water} = 293.150 \cdot K$

$T_{final} := (273.15 + 27.70) \cdot K$ $\quad$ $T_{final} = 300.850 \cdot K$

First determine how much energy is gained by the water when the piece of titanium is added. We know how much water there is and the temperature change of the water. The specific heat of water is well known (and given in your textbook). From the specific heat and the amount of water, determine the heat capacity. From the temperature change and the heat capacity, determine the amount of energy transfered.

$C_{water} := 4.184 \cdot joule \cdot gm^{-1} \cdot K^{-1}$

This is the specific heat of water. Notice the units. This is the amount of energy (4.184 joule) required to change the temperature of one gram of water by 1 kelvin.

$density_{water} := 0.99823 \cdot gm \cdot mL^{-1}$

Value from the CRC Handbook of Chemistry and Physics. If you look in the Index, you can find a value for water at 25 ºC. Or you can assume that the density of water is approximately 1.0 g mL^{-1}.

$mass = density \cdot volume$

You should know this. Think about what density means and you should be able to figure out the equation. Also, double check yourself by thinking about the units.

$mass_{water} := density_{water} \cdot volume_{water}$

$mass_{water} = 998.230 \cdot gm$

Watch your units. You need to convert liters to mL, then multiply by the density to get the mass of the water.

$HC_{water} := mass_{water} \cdot C_{water}$

$HC_{water} = 4.177 \cdot 10^{3} \cdot joule \cdot K^{-1}$

The heat capacity of the water. This is the amount of energy (joules) required to change the water temperature by 1 Kelvin.

$\Delta T_{water} := T_{final} - T_{water}$

Dela (Δ) is the symbol used for change. So this is the change in the temperature of the water.

$energy := HC_{water} \cdot \Delta T_{water}$

$energy = 3.216 \cdot 10^{4} \cdot joule$

The energy required to heat the water from the initial to the final temperature. The key to this problem is recognizing that the energy gained by the water (to heat it) all comes from the piece of hot titanium. So the energy gained by the water is equal to the energy lost by the titanium.

Now that we know the energy lost by the piece of titanium metal. This is used to calculate the specific heat of titanium.

$\Delta T_{metal} := T_{metal} - T_{final}$

The temperature change of the piece of titanium.

$\Delta T_{metal} = 72.300 \cdot K$

$energy = HC \cdot \Delta T$ has solution(s) $HC = \frac{energy}{\Delta T}$

This demonstrates how Mathcad preforms algebra to rearange an equation. This takes a relationship (equataion) that was developed earlier and rearanges it to find the variable we are currently interested in.

$HC_{metal} := \frac{energy}{\Delta T_{metal}}$

Calculate the heat capacity for the peice of titanium.

$HC_{metal} = 444.810 \cdot joule \cdot K^{-1}$

$HC = mass \cdot C$

$C_{metal} := \frac{HC_{metal}}{mass_{metal}}$

Calculate the specific heat capacity for titanium.

$C_{metal} = 0.522 \cdot joule \cdot gm^{-1} \cdot K^{-1}$

This compares to the value of 0.125 cal g^{-1} K^{-1} in the *CRC Handbook of Chemisty and Physics*. Newer editions of the handbook will have SI units

$C_{metal} = 0.125 \cdot cal \cdot gm^{-1} \cdot K^{-1}$

Next the same piece of titanium is heated in an acetylene flame (like that used for welding) to an unknown temperature. When the piece of titanium is placed in a 10.000 liter container of water at 20.00 ºC the final temperature is now 32.72 ºC. What is the temperature of the flame? At what temperature does titanium melt?

$T_{water} := (273.15 + 20.00) \cdot K$ $\qquad T_{water} = 293.150 \cdot K$

$T_{final} := (273.15 + 32.72) \cdot K$ $\qquad T_{final} = 305.870 \cdot K$

$volume_{water} := 10.0 \cdot liter$

Calculate the change in temperature:

$\Delta T_{water} := T_{final} - T_{water}$ $\qquad \Delta T_{water} = 12.720 \cdot K$

Calculate the mass of the water:

$mass_{water} := density_{water} \cdot volume_{water}$ $\qquad mass_{water} = 9.982 \cdot 10^{3} \cdot gm$

Calculate the heat capacity of the water:

$HC_{water} := mass_{water} \cdot C_{water}$ $\qquad HC_{water} = 4.177 \cdot 10^{4} \cdot joule \cdot K^{-1}$

Calculate the energy gained by the water:

$energy := HC_{water} \cdot \Delta T_{water}$ $\qquad energy = 5.313 \cdot 10^{5} \cdot joule$

Since the energy gained by the water must have been lost by the metal. Using the heat capacity for this peice of metal (determined above).

$HC_{metal} = 444.810 \cdot joule \cdot K^{-1}$

$energy = HC_{metal} \cdot \Delta T_{metal}$

$\Delta T_{metal} := \dfrac{energy}{HC_{metal}}$ $\qquad \Delta T_{metal} = 1.194 \cdot 10^{3} \cdot K$

The initial temperature of the metal:

$T_{metal} := T_{final} + \Delta T_{metal}$ $\qquad T_{metal} = 1.500 \cdot 10^{3} \cdot K$

Look up the melting point of Titanium in the *CRC Handbook of Chemistry and Physics*. Should the piece of titanum melt at this temperature?

3. Calculate the energy required to heat a 55.4 g ice cube that starts in a freezer at -100.0 ^{o}C (VERY COLD):

a. Heat from the freezer to ice at 0.0 ^{o}C

Information given in the problem:

$$mass := 55.4 \cdot gm$$

$$T_{initial} := (273.15 - 100) \cdot K \qquad T_{initial} = 173.150 \cdot K$$

$$T_{final} := (273.15 + 0) \cdot K \qquad T_{final} = 273.150 \cdot K$$

The temperature change:

$$\Delta T := T_{final} - T_{initial} \qquad \Delta T = 100.000 \cdot K$$

Energy required:

$$C_{ice} := 2.1 \cdot joule \cdot gm^{-1} \cdot K^{-1}$$

(The specific heat of ice. Notice that it is different that the specific heat of liquid water or steam. Also notice the units.)

$$energy_{ice} := C_{ice} \cdot mass \cdot \Delta T$$

$$energy_{ice} = 1.163 \cdot 10^{4} \cdot joule$$

b. Heat the ice at 0.0 ^{o}C to liquid at 0.0 ^{o}C.

$$\Delta H_{fusion} := 333 \cdot joule \cdot gm^{-1}$$

$$energy_{melt} := \Delta H_{fusion} \cdot mass$$

$$energy_{melt} = 1.845 \cdot 10^{4} \cdot joule$$

(The heat of fusion (ΔH_{fusion}) is the amount of energy required to melt or the amout of energy released durring freezing. Compare the energy required to melt the ice with the energy required to heat it in part a.)

c. Heat from liquid at 0.0 ^oC to liquid at 100.0 ^oC.

Information given in the problem:

$T_{initial} := (273.15 - 0.0) \cdot K$ $\quad$ $T_{initial} = 273.150 \cdot K$

$T_{final} := (273.15 + 100.0) \cdot K$ $\quad$ $T_{final} = 373.150 \cdot K$

The temperature change:

$\Delta T := T_{final} - T_{initial}$ $\quad$ $\Delta T = 100.000 \cdot K$

Energy required:

$C_{water} := 4.2 \cdot joule \cdot gm^{-1} \cdot K^{-1}$

$energy_{water} := C_{water} \cdot mass \cdot \Delta T$

$energy_{water} = 2.327 \cdot 10^4 \cdot joule$

d. Heat the liquid at 100.0 ^oC to gas at 100.0 ^oC.

$\Delta H_{vaporization} := 2260 \cdot joule \cdot gm^{-1}$

$energy_{vaporization} := \Delta H_{vaporization} \cdot mass$

$energy_{vaporization} = 1.252 \cdot 10^5 \cdot joule$

(The heat of vaporization ($\Delta H_{vaporization}$) is the amount of energy required to boil or the amout of energy released durring condensation. Compare this energy to that required for all the previous steps combined. Why are steam burns so bad?)

e. Heat from gas at 100.0 ^oC to gas at 200.0 ^oC.

Information given in the problem:

$T_{initial} := (273.15 + 100.0) \cdot K$ $\quad$ $T_{initial} = 373.150 \cdot K$

$T_{final} := (273.15 + 200.0) \cdot K$ $\quad$ $T_{final} = 473.150 \cdot K$

The temperature change:

$\Delta T := T_{final} - T_{initial}$ $\quad$ $\Delta T = 100.000 \cdot K$

Energy required:

$C_{steam} := 2.0 \cdot joule \cdot gm^{-1} \cdot K^{-1}$

$energy_{steam} := C_{steam} \cdot mass \cdot \Delta T$

$energy_{steam} = 1.108 \cdot 10^4 \cdot joule$

f. Heat from ice at -100.0 ^oC to gas at 200.0 ^oC.

$$energy_{total} := energy_{ice} + energy_{melt} + energy_{water} + energy_{vaporization} + energy_{steam}$$

$$energy_{total} = 1.896 \cdot 10^5 \cdot joule$$

4. Balance the following reactions, calculate ΔH_{rxn}, calculate the energy released (or required) for the combustion of 1.50 grams of fuel (H_2, CH_4 (methane), or C_4H_{10} (butane) in each reaction and compare the energetics for the reactions (which is the best fuel?).

a. $2\ H_2\ (g)\ +\ 1\ O_2\ (g)\ \text{-->}\ 2\ H_2O\ (g)$

$$kJ := 10^3 \cdot joule$$

$\Delta H_{formation}$ for the products and reactants (from table in textbook):

$$\Delta H_{H2Og} := -241.8 \cdot kJ \cdot mole^{-1}$$

$$\Delta H_{H2} := 0 \cdot kJ \cdot mole^{-1}$$

$$\Delta H_{O2} := 0 \cdot kJ \cdot mole^{-1}$$

Calculate ΔH_{rxn}

$$\Delta H_{rxn} := \left(2 \cdot \Delta H_{H2Og}\right) - \left[\left(1 \cdot \Delta H_{H2}\right) + \left(2 \cdot \Delta H_{O2}\right)\right]$$

$$\Delta H_{rxn} = -483.600 \cdot kJ \cdot mole^{-1}$$

Energy for reaction as written. With 2 moles of H_2 reacting with 1 mole of O_2 to produce 2 moles of H_2O

Calculate number of moles for 1.50 grams of H_2

$$H2_{mass} := 1.50 \cdot gm$$

$$H2_{MW} := (2 \cdot 1.00794) \cdot gm \cdot mole^{-1} \qquad H2_{MW} = 2.016 \cdot gm \cdot mole^{-1}$$

$$H2 := \frac{H2_{mass}}{H2_{MW}} \qquad H2 = 0.744 \cdot mole$$

Calculate energy for the reaction of 1.50 grams of H_2 in excess oxygen

$$energy := \Delta H_{rxn} \cdot \frac{H2}{2 \cdot mole} \qquad energy = -179.921 \cdot kJ$$

b. $2\ CH_4\ (g) + 3\ O_2\ (g) \rightarrow 2\ CO\ (g) + 4\ H_2O\ (g)$

$\Delta H_{formation}$ for the products and reactants (from table in textbook):

$$\Delta H_{H2Og} := -241.8 \cdot kJ \cdot mole^{-1}$$

$$\Delta H_{CO} := -110.5 \cdot kJ \cdot mole^{-1}$$

$$\Delta H_{O2} := 0 \cdot kJ \cdot mole^{-1}$$

$$\Delta H_{CH4} := -74.8 \cdot kJ \cdot mole^{-1}$$

Calculate ΔH_{rxn}

$$\Delta H_{rxn} := \left[\left(4 \cdot \Delta H_{H2Og}\right) + \left(2 \cdot \Delta H_{CO}\right)\right] - \left[\left(2 \cdot \Delta H_{CH4}\right) + \left(3 \cdot \Delta H_{O2}\right)\right]$$

$$\Delta H_{rxn} = -1.039 \cdot 10^{3} \cdot kJ$$

Energy for reaction as written. With 2 moles of CH_4 reacting with 3 mole of O_2 to produce 2 moles of CO and 4 moles of H_2O

Calculate number of moles for 1.50 grams of CH_4

$$CH4_{mass} := 1.50 \cdot gm$$

$$CH4_{MW} := (12.011 + (4 \cdot 1.00794)) \cdot gm \cdot mole^{-1} \qquad CH4_{MW} = 16.043 \cdot gm \cdot mole^{-1}$$

$$CH4 := \frac{CH4_{mass}}{CH4_{MW}} \qquad CH4 = 0.094 \cdot mole$$

Calculate energy for the reaction of 1.50 grams of CH_4 in excess oxygen

$$energy := \Delta H_{rxn} \cdot \frac{CH4}{2 \cdot mole}$$

$$energy = -48.555 \cdot kJ$$

c. 1 CH_4 (g) + 2 O_2 (g) --> 1 CO_2 (g) + 2 H_2O (g)

$\Delta H_{formation}$ for the products and reactants (from table in textbook):

$$\Delta H_{H2Og} := -241.8 \cdot kJ \cdot mole^{-1}$$

$$\Delta H_{CO2} := -393.5 \cdot kJ \cdot mole^{-1}$$

$$\Delta H_{O2} := 0 \cdot kJ \cdot mole^{-1}$$

$$\Delta H_{CH4} := -74.8 \cdot kJ \cdot mole^{-1}$$

Calculate ΔH_{rxn}

$$\Delta H_{rxn} := \left[\left(2 \cdot \Delta H_{H2Og}\right) + \left(1 \cdot \Delta H_{CO2}\right) \right] - \left[\left(1 \cdot \Delta H_{CH4}\right) + \left(2 \cdot \Delta H_{O2}\right) \right]$$

$$\Delta H_{rxn} = -802.300 \cdot kJ$$

Energy for reaction as written. With 1 mole of CH_4 reacting with 2 mole of O_2 to produce 1 moles of CO_2 and 2 moles of H_2O

Calculate number of moles for 1.50 grams of CH_4

$$CH4_{mass} := 1.50 \cdot gm$$

$$CH4_{MW} := (12.011 + (4 \cdot 1.00794)) \cdot gm \cdot mole^{-1} \qquad CH4_{MW} = 16.043 \cdot gm \cdot mole^{-1}$$

$$CH4 := \frac{CH4_{mass}}{CH4_{MW}} \qquad CH4 = 0.094 \cdot mole$$

Calculate energy for the reaction of 1.50 grams of CH_4 in excess oxygen

$$energy := \Delta H_{rxn} \cdot \frac{CH4}{1 \cdot mole}$$

$$energy = -75.015 \cdot kJ$$

Notice that more energy is released when the methane reacts to produce CO_2 compared with the reaction that produces CO.

d. 2 C_4H_{10} (g) + 13 O_2 (g) --> 8 CO_2 (g) + 10 H_2O (g)

$\Delta H_{formation}$ for the products and reactants (from table in textbook):

$$\Delta H_{H2Og} := -241.8 \cdot kJ \cdot mole^{-1}$$

$$\Delta H_{CO2} := -393.5 \cdot kJ \cdot mole^{-1}$$

$$\Delta H_{O2} := 0 \cdot kJ \cdot mole^{-1}$$

$$\Delta H_{C4H10} := -888.0 \cdot kJ \cdot mole^{-1}$$

Calculate ΔH_{rxn}

$$\Delta H_{rxn} := \left[\left(10 \cdot \Delta H_{H2Og}\right) + \left(8 \cdot \Delta H_{CO2}\right) \right] - \left[\left(2 \cdot \Delta H_{C4H10}\right) + \left(13 \cdot \Delta H_{O2}\right) \right]$$

$$\Delta H_{rxn} = -3.790 \cdot 10^3 \cdot kJ$$

Energy for reaction as written. With 2 mole of C_4H_{10} reacting with 13 mole of O_2 to produce 8 moles of CO_2 and 10 moles of H_2O

Calculate number of moles for 1.50 grams of C_4H_{10}

$$C4H10_{mass} := 1.50 \cdot gm$$

$$C4H10_{MW} := ((4 \cdot 12.011) + (10 \cdot 1.00794)) \cdot gm \cdot mole^{-1} \qquad C4H10_{MW} = 58.123 \cdot gm \cdot mole^{-1}$$

$$C4H10 := \frac{C4H10_{mass}}{C4H10_{MW}} \qquad C4H10 = 0.026 \cdot mole$$

Calculate energy for the reaction of 1.50 grams of C_4H_{10} in excess oxygen

$$energy := \Delta H_{rxn} \cdot \frac{C4H10}{2 \cdot mole} \qquad energy = -48.905 \cdot kJ$$

Based on the amount of energy released for a given mass of fuel, hydrogen is the best. Reactions b and c show that combustion without sufficent oxygen (reaction b) is much less efficient.

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