

Practice sheet #8: thermodynamics.

1. We measure heat in Joules because it is a form of energy; but another unit for energy is the calorie. The calorie is defined as the amount of energy necessary to increase the temperature of 1 gram of liquid water by 1 °C. Knowing that the molar heat capacity for water is 75.24 J/K mol and using dimensional analysis, calculate the conversion factor between calorie and joule.
2. When a ball of mass 5.6Kg is dropped through a height, its potential energy changes. Suppose that, when the ball hits the ground, all its energy is converted into heat, increasing the temperature of the ball. If the specific heat capacity of the material of the ball is $0.850 \text{ J K}^{-1} \text{ g}^{-1}$, from what height do we need to drop the ball to increase its temperature by 1.00°C? (Hint: the gravitation potential energy is $E_p = mgh$)

3. During the process to prepare ammonia (NH_3), nitrogen (N_2), is heated slowly, maintaining the external pressure close to the internal one of 50.0 atm, until its volume has increased from 542 L to 974 L. Calculate the work done on the nitrogen as it is heated.
4. A cylinder confines 2.00 L of gas under a pressure of 1.00 atm. The external pressure is also 1.00 atm. The gas is heated slowly, with the piston sliding freely to maintain the pressure of the gas close to 1.00 atm. The heating continues until the volume is 3.50 L. Calculate the work done on the gas.

5. In the “Memoir on heat”, published in 1783, Lavoisier and Laplace reported, “The heat necessary to melt ice is equal to three quarters of the heat that can raise the temperature of the same amount from its melting to its boiling point.”. Knowing that the heat capacity of liquid water is $4.18 \text{ J/}^\circ\text{Cg}$, use this observation to calculate how much heat is needed to melt 1.00 g of ice.

6. A calorimeter is calibrated using the following experimental data:

Heat (KJ)	ΔT ($^\circ\text{C}$)	Heat (KJ)	ΔT ($^\circ\text{C}$)
5.00	2.17	20.0	8.10
10.0	4.12	25.0	11.50
15.0	6.98	30.0	11.06

After calibration, that same calorimeter was used to calculate the heat released by the combustion of an unknown hydrocarbon. Referring to the table on page A17 of the book (Appendix 2A/Organic Compounds/Hydrocarbons), and knowing that the combustion of 10.00 mol of this unknown gas cause the temperature in the calorimeter to raise from 70.0 F to 104.2 F, are you able to identify the hydrocarbon?

7. Enthalpy of physical changes:

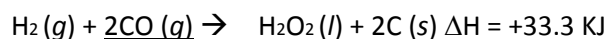
- a. Steam (water vapor) at 167 °C is bubbled in liquid water at 25 °C to quickly heat it up. Calculate how much steam is needed to bring to ebullition 1.24 Kg of water.

- b. Calculate how much steam would be needed if instead of liquid water, you were to use 1.24Kg of solid ice at -18 °C.

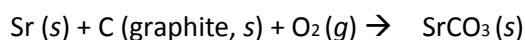
- c. Once the water from question 3b is boiling, how much more ice at -18 °C would be needed to freeze the overall amount of water?

8. Enthalpy of chemical changes:

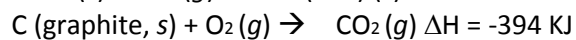
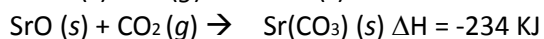
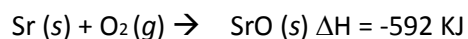
- a. For each of the following reaction, the ΔH is the enthalpy change when the number of moles indicated by the stoichiometric coefficients react. Calculate the ΔH when 1.00 g of the underlined substance is consumed or produced:



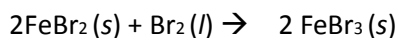
- b. Calculate the enthalpy change for the formation of 1.00 mol of strontium carbonate (the material used to obtain red colored fireworks) from its elements. Note: all reactions need to be balanced.



Available information is:



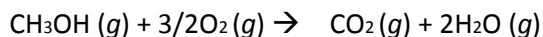
- d. The molar enthalpy of formation for FeBr_2 is -249.8 kJ/mol and that for FeBr_3 is -268.2 kJ/mol . What is the ΔH of the following reaction?



9. Summary problem: **Methanol as a gasoline substitute.** *This problem is designed to summarize the whole Chapter, if you can solve it, you have properly understood and mastered chapter 6..*

Methanol (CH_3OH) is used as a substitute for gasoline in certain high-performance vehicles. To design engines that will run on it, its thermochemistry must be understood.

- The methanol in an automobile engine must be in the gas phase before it can react. Calculate the heat (in kJ) that must be added to 1.00 Kg of liquid methanol to raise its temperature from 25°C to its normal boiling point, 65°C , knowing that the molar heat capacity of liquid methanol is $81.6 \text{ J/}^\circ\text{Cmol}$.
- Once the methanol has reached its boiling point, it must be vaporized. The molar enthalpy of vaporization for methanol is 38 kJ/mol . How much heat must be added to vaporize 1.00 Kg of methanol?
- Once it is in the vapor phase, the methanol can react with oxygen in the air according to



Use average bond enthalpies to estimate the enthalpy change in this reaction, for one mole of methanol reacting.

- Use data from Appendix 2A of your textbook to calculate the actual enthalpy change in this reaction, assuming it to be the same at 65°C as at 25°C .
- Calculate the heat released when 1.00 Kg of gaseous methanol is burned in air at constant pressure. Use result of part d.
- Calculate the difference between the change in enthalpy and the change in internal energy when 1.00 Kg of gaseous methanol is oxidized to gaseous CO_2 and H_2O at 65°C .
- Suppose now that the methanol is burned inside the cylinder of an automobile. Taking the radius of the cylinder to be 4.0 cm and the distance moved by the piston during one stroke to be 12.0 cm, calculate the work done on the gas per stroke as it expands against an external pressure of 1.00 atm. Express your answer in both, L atm and J.