

Strand D1: THERMODYNAMICS: Energy Changes in Chemical Reactions

(Nelson text: Unit 3 Chapter 5 - pg 294 - 356 & Chapter 7 (7.7) page 494-508)

D1:1. Write a balanced equation for the combustion of ethanol (C₂H₅OH).

What would be a third product of this reaction ?

Calorimetry

D1:2. Attempt question # 8, 9, 10, & 11 on page 302 in your text

D1:3. If a concrete wall of mass 4.5×10^7 g reached a temperature of 38°C during the day, how much energy would it release if its temp dropped to 18°C that night ? (c = 3.340 kJ.kg⁻¹K⁻¹ for concrete)

D1:4. Fat tissue is 85% fat and 15% water. When fat is “burned” in the body it releases **9.0 kcal per gram** of fat.

a) How many kilocalories are released by the loss of 1.0 lb of fat on your “diet” program

b) Running **15.0 km/h** expends about **500 kcal/h** of energy. How far does a person have to run to burn off the 1.0 lb of fat tissue by this means only ?

Enthalpy Changes (5.3 pg 313-319)

D1:5. Attempt questions # 1, 2, & 3 on page 319 in your text.

D1:6. Given $3 \text{ C (s)} + 2 \text{ Fe}_2\text{O}_3 \text{ (s)} + 462 \text{ kJ} \rightarrow 4 \text{ Fe (s)} + 3 \text{ CO}_2 \text{ (g)}$

Rewrite this equation using **1 mole** of carbon and ΔH notation

D1:7. Attempt questions # 2 on page 320 in your text.

D1:8. Calcium is burnt according to the following eq'n $\text{Ca (s)} + 1/2 \text{ O}_2 \text{ (g)} \rightarrow \text{CaO (s)} \quad \Delta H = -636 \text{ kJ/mol Ca}$

How much heat would be produced by burning **12 g** of calcium ?

D1:9. a. Magnesium burns in air according to the following thermochemical equation to produce a bright light and is often used in fireworks displays. $2 \text{ Mg (s)} + \text{ O}_2 \text{ (g)} \rightarrow 2 \text{ MgO (s)} \quad \Delta H^\circ = -1203 \text{ kJ}$

How much heat is liberated by the combustion of **6.54 g** of magnesium?

b. Methanol is the fuel in “canned heat” or sterno cans that are used to heat foods at buffet tables. The heat of combustion for methanol is 600 kJ / mol. How much heat is produced by combusting a 46 g can of sterno?

Standard Enthalpies of Formation - ΔH°_f (5.5 pg 331-332)

D1:10. Attempt question # 1 on page 332 in your text.

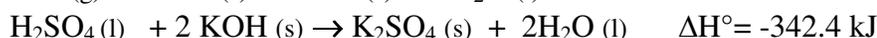
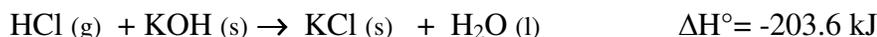
D1:11. Write the equation for the formation of the following, include ΔH°_f .

a) HC₂H₃O₂ b) H₂SO₄ c) NaNO₃ d) Ba(ClO₃)₂

D1:12. Calculate ΔH_{rx} for : a) $2 \text{ Fe (s)} + \text{ O}_2 \text{ (g)} \rightarrow 2 \text{ FeO (s)}$ b) $2 \text{ NH}_3 \text{ (g)} \rightarrow \text{ N}_2 \text{ (g)} + 3 \text{ H}_2 \text{ (g)}$

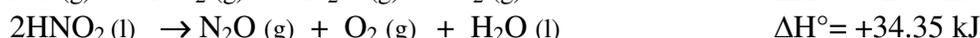
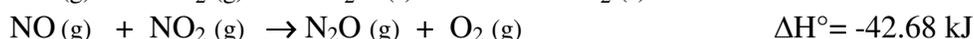
Hess's Law of Heat Summation (5.4 pg 322-329)

D1:13. Calculate ΔH° for the following - $2 \text{ KCl (s)} + \text{ H}_2\text{SO}_4 \text{ (l)} \rightarrow 2 \text{ HCl (g)} + \text{ K}_2\text{SO}_4 \text{ (s)}$
using:



D1:14. Calculate ΔH° for the following - $\text{HCl (g)} + \text{NaNO}_2 \text{ (s)} \rightarrow \text{HNO}_2 \text{ (l)} + \text{NaCl (s)}$

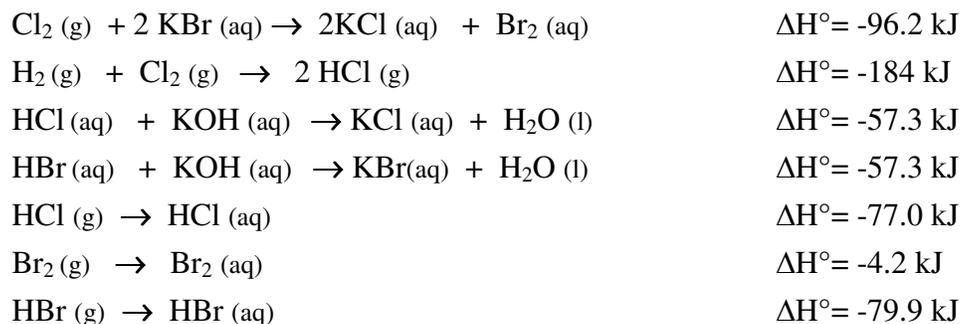
using:



*D1:15. Attempt question # 1, 2 & 3 on page 326 in your text.

*D1:16. Attempt question # 5 on page 329 in your text.

*D1:17. Calculate ΔH°_f for HBr (g) using the following experimental data



Calculating ΔH_{rx} using the Σ method (5.5 pg 332-335)

D1:18. Attempt question # 2 & #5 on page 335 in your text.

D1:19. Calculate ΔH_{rx} for the following a) $2 \text{NH}_3(\text{g}) + \text{CO}_2(\text{g}) \rightarrow \text{CO}(\text{NH}_2)_2(\text{s}) (\text{urea}) + \text{H}_2\text{O}(\text{l})$

D1:20. Attempt question # 3 on page 335 in your text.

D1:21. The standard heat of combustion (ΔH_{rx}) for eicosane ($\text{C}_{20}\text{H}_{42}$), a typical component of candle wax, is **47.3 kJ/g** when it burns in pure oxygen. The only products are $\text{CO}_2(\text{g})$ and $\text{H}_2\text{O}(\text{g})$. Calculate the standard heat of formation (ΔH°_f) of eicosane in kJ/mol.

D1:22. **3.0 g** of hydrazine (N_2H_4) are burned in a bomb calorimeter causing **3.9 L** of water to increase its temperature by **3.5°C**.

Calculate the heat of combustion (ΔH_{comb}) for hydrazine in kJ/g and kJ/mol.

Connections

D1:23. Attempt question # 4 on page 340 in your text.

D1:24. a) Write the equation for the complete combustion of benzene (C_6H_6 (P)).

b) **2.1 g** of benzene is burned in a metal bomb calorimeter (*mass = 80 g, c = 0.825 kJ/kgK*) containing **3800 mL** of water. The temperature rose from **22°C** to **27.2°C**. Calculate the **experimental** heat of combustion (ΔH_{comb}) for benzene with this data.

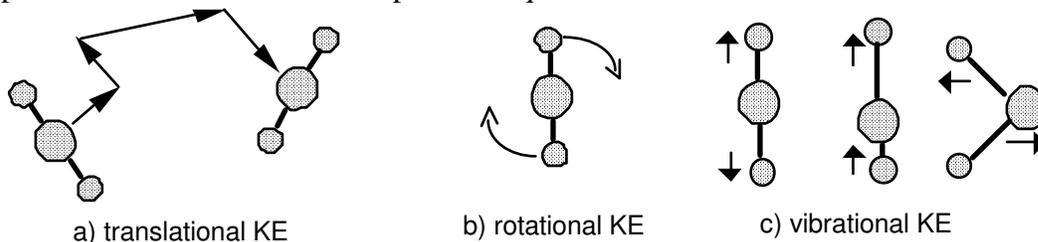
c) Now use your data sheet to calculate the **theoretical** $\Delta H_{(\text{combustion})}$ for benzene

d) Calculate the % error in the experiment and do a short error analysis.

***for your notes* - The origins of ΔH**

All molecules possess both KINETIC and POTENTIAL energy !!

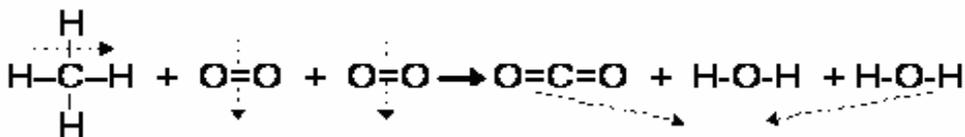
Kinetic energy is due to motion and all molecules above 0 Kelvin are moving thus have some KE (see diagram below). Solids would have less than liquids and liquids less than gases. Relatively, the total value of energies tied up in these motions is small - esp solids/liquids.



Potential energy is due to the relative position of atoms to each other i.e. the bonding patterns that the molecules or atom groups possess. Interatomic forces could create either ionic bonds or covalent bonds whereas intermolecular forces are created when molecules form aggregates (solids) i.e. hydrogen bonds, vanderWaal bonds etc. The total value of energy tied up in these relationships tends to be relatively large.

Since most of a molecule's internal energy is PE (bond energy) then it follows that when a molecule undergoes change (reacts to form a new molecule by breaking its original bonds (energy required) and reforming new bonds (energy released) the resulting net change in enthalpy (ΔH_{rx}) that you the chemist observes is primarily a change in Potential energy - so $\Delta H_{rx} = \Delta PE$

See example over



energy **required** to break these bonds ($\Delta H=+ve$)

energy **released** when these bonds are formed ($\Delta H=-ve$)

$\Delta H_{rx} = \text{net change in PE}$

Bond Energies and Heats of Reaction

- D1:25. a) What energy would be required to completely disassemble a mole of butane gas (C_4H_{10}) molecules into individual gaseous atoms. (= *total bonding energy*) *hint: draw the structural diagram!*
 b) How much energy would be released when 1 mole of acetone (CH_3COCH_3) is formed?

*D1.26 The standard heat of formation (ΔH°_f) for $SF_6(g)$ is **-1096 kJ/mol**. The energy required to gasify solid sulfur is **277 kJ/mol**. Calculate the "average" S—F bond energy in SF_6 .

Spontaneity / Entropy / Gibbs Free Energy (7.7 pg 494-508)

D1:27 According to the 1st Law of Thermodynamics **only**, is the following reaction spontaneous
 $C_2H_2(g) + 5 N_2O(g) \rightarrow 2 CO_2(g) + H_2O(g) + 5 N_2(g)$

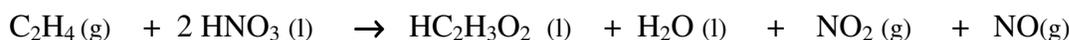
D1:28. *Entropy predictions*-Do # 1 on page 498 & #3 + #10 on pg 512 & #19 + #20 on pg 522 in your text.

D1:29 According to the 2nd Law of Thermodynamics **only**, are the following reactions spontaneous (ie find ΔS)

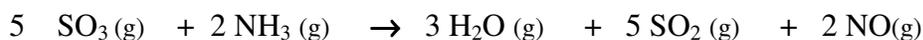
- a) $N_2(g) + 3 H_2(g) \rightarrow 2 NH_3(g)$
 b) $S(s) + 2 N_2O(g) \rightarrow SO_2(g) + 2 N_2(g)$

D1:30. *Gibbs Free Energy*-Do # 2 c on page 508 & #11 b, + #14 on page 512 in your text

D1:31 Determine whether the following reaction will be spontaneous at 25°C.



D1:32 Determine whether the following reaction will be spontaneous at 100°C.



*1:33. Given: $Cu_2O(s) + 0.5 O_2(g) \rightarrow 2 CuO(s)$

Calculate ΔG° for this rx then discuss the 2 laws of Thermodynamics as they apply to this reaction.

Determine the temperature at which there will be no reaction (equilibrium) and discuss the spontaneity of the reaction above and below this temperature (ie the temperature range of this rx)

Free Energy and equilibrium (ie $\Delta G = 0$)

D1:34. Liquid chloroform ($CHCl_3$) has a heat of vapourization of +31.4 kJ/mol ie $CHCl_3(l) \rightarrow CHCl_3(g)$ and an entropy increase of 94.2 J/mol.K. At what temperature do we expect chloroform to boil (ie liquid & vapour in equilibrium at 101.3 kPa)

D1:35 Acetone has a boiling point of 56.2°C. The change $(CH_3)_2CO(l) \rightarrow (CH_3)_2CO(g)$ has a $\Delta H^\circ = 31.9$ kJ/mol. What is the ΔS for this change?

*D1:36. Microwaves are used to heat food. The microwave radiation is absorbed by moisture in the food. This heats the water and thus the food. How many photons having a wavelength of **3.0 mm** would have to be absorbed by **1.0 g** of water to raise its temperature by **1°C**.

(HINT: Find Q (req'd) then E per microwave using

$E_{\text{photon}} = hc/\lambda$ where h (planck's constant) = 6.6×10^{-34} J.s & c = speed of light = 3.0×10^8 m/s

Assignment D1: Thermodynamics

Total = ___/25 marks

These are sample questions similar to “test” questions that you should do to see if you understand the concepts in this unit. Copying them from elsewhere defeats this purpose!

They should be handed in on the day you write the unit test AT THE BEGINNING OF THE CLASS. Any other time will be considered “late” and deductions will occur

1. Pentane is a hydrocarbon that is commonly used as a fuel in lighters. [7 marks]
 - a. Write a balanced equation for the complete combustion of pentane gas.
 - b. Calculate i) the volume of oxygen at 30°C and 100 kPa & ii) the volume of “air” at 30°C and 100 kPa that is needed to completely burn 10.0 g of pentane.
 - c. The *heat of combustion* (ΔH_{comb}) of pentane is -3220 kJ/mol. Use the Σ *method* to calculate the heat of formation ($\Delta H^{\circ}_{\text{f}}$) of pentane.
 - d. Assuming that all of the heat evolved in burning 30.0 g of pentane is transferred to 8.0 kg of water, how much will the temperature of the water change by?
2. Ethyne gas (C_2H_2) can be “hydrogenated” to produce ethane gas. [7 marks]
 - a. Write a balanced equation for this hydrogenation reaction.
 - b. Using your “thermochemical” tables calculate the value of ΔG° for this addition reaction and comment on its spontaneity at this temperature.
 - c. You know that ΔH is energy that originates in “bonds”-breaking bonds requires energy, making bonds releases energy.
Now use your “bond dissociation energy” tables (3&4) and the ΔH° value from “b” to estimate the bond dissociation energy for the $\text{C}\equiv\text{C}$ bond in ethyne.
Compare this calculated value to the actual value in table 4 and explain the discrepancy.
3. Gaseous hydrogen sulfide will react with sulfur dioxide to produce solid sulfur and water vapour in the large INCO stack in Sudbury [6 marks]
 - a. Write a balanced equation for this reaction.
 - b. Determine if this reaction will occur in the smoke stack which is at 425°C.
 - c. Write a “technical” note to INCO explaining to them how to prevent this reaction from happening in the stack.
4. Phosphorus trichloride will burn in pure chlorine gas to produce phosphorus pentachloride. 15.0 g of phosphorus trichloride were combusted in a metal calorimeter and all the heat generated collected in a 50 mL water initially at 23°C. When the reaction was over the water temperature was 64°C. [5 marks]
 - a. Calculate the % error for this experiment
 - b. Do an error analysis (ie how could you reduce the error)