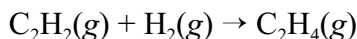


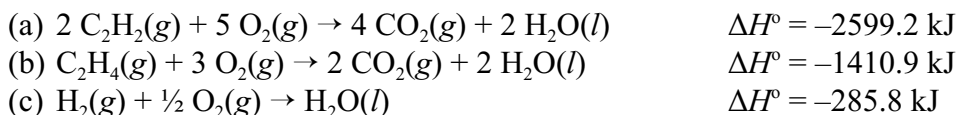
Chem 115 POGIL Worksheet - Week 8 - Answers
Thermochemistry (Continued), Electromagnetic Radiation, and Line Spectra

Key Questions & Exercises

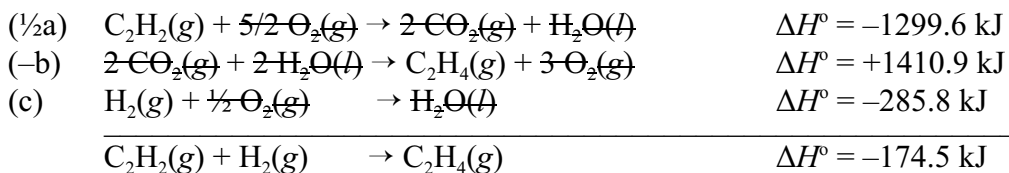
1. Calculate ΔH° for the reaction,



Given:



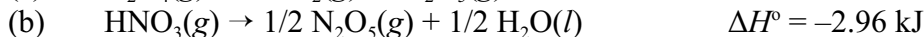
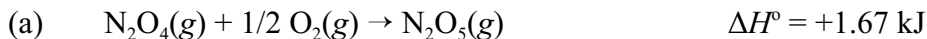
The (a) equation has $\text{C}_2\text{H}_2(\text{g})$ on the left, where we need it in the target equation, but its coefficient is 2. If we take one half of equation (a) we will have the correct coefficient. The ΔH° contribution will be $(\frac{1}{2})(-2599.2 \text{ kJ}) = -1299.6 \text{ kJ}$. Equation (b) has $\text{C}_2\text{H}_4(\text{g})$ with the correct coefficient for our target equation, but we need it on the right, not the left as given. Therefore, we will reverse it, making its ΔH° contribution be $+1410.9 \text{ kJ}$. Finally, if we take equation (c) as given, it will introduce the $\text{H}_2(\text{g})$ with the correct coefficient, and it will result in cancellations of the $\text{CO}_2(\text{g})$, $\text{H}_2\text{O}(\text{l})$, and $\text{O}_2(\text{g})$ introduced by our manipulations of the first two equations, which are absent in the target equation. The following sum results:



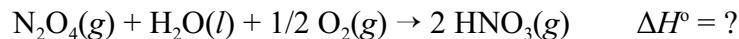
2. Write the balanced thermochemical equations that pertain to the standard enthalpies of formation of the given compounds.

Compound	ΔH_f° (kJ/mol)	Thermochemical Equation
$\text{CCl}_4(\text{g})$	-106.7	$\text{C}(\text{s}) + 2 \text{Cl}_2(\text{g}) \rightarrow \text{CCl}_4(\text{l})$
$\text{Fe}_2\text{O}_3(\text{s})$	-822.16	$2 \text{Fe}(\text{s}) + \frac{3}{2} \text{O}_2(\text{g}) \rightarrow \text{Fe}_2\text{O}_3(\text{s})$
$\text{HNO}_3(\text{g})$	-134.3	$\frac{1}{2} \text{H}_2(\text{g}) + \frac{1}{2} \text{N}_2(\text{g}) + \frac{3}{2} \text{O}_2(\text{g}) \rightarrow \text{HNO}_3(\text{g})$
$\text{NaHCO}_3(\text{s})$	-947.7	$\text{Na}(\text{s}) + \frac{1}{2} \text{H}_2(\text{g}) + \text{C}(\text{s}) + \frac{3}{2} \text{O}_2(\text{g}) \rightarrow \text{NaHCO}_3(\text{s})$

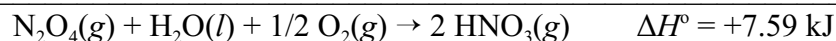
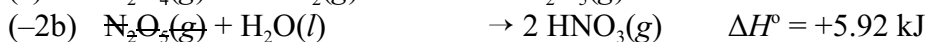
3. i. Given:



Calculate ΔH° for the reaction



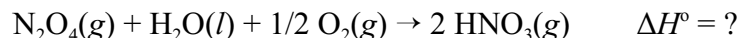
The two given equations, manipulated as shown below, will add to give the target reaction and its enthalpy:



ii. Given following standard enthalpy of formation data:

Compound	ΔH_f° (kJ/mol)
$\text{N}_2\text{O}_4(g)$	+9.66
$\text{HNO}_3(g)$	-134.31
$\text{H}_2\text{O}(l)$	-285.83

Calculate ΔH° for the reaction



Compare your answer to your answer in part i.

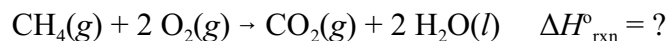
The enthalpy for the reaction is the sum of the enthalpies of formation for products minus those for reactants, each multiplied by its stoichiometric coefficient in the balanced equation:

$$\begin{aligned} \Delta H^\circ &= 2\Delta H_f^\circ(\text{HNO}_3) - \{\Delta H_f^\circ(\text{N}_2\text{O}_4) + \Delta H_f^\circ(\text{H}_2\text{O}) + 1/2\Delta H_f^\circ(\text{O}_2)\} \\ &= (2)(-134.31 \text{ kJ}) - \{+9.66 \text{ kJ} + (-285.83 \text{ kJ}) + 0 \text{ kJ}\} \\ &= 7.55 \text{ kJ} \end{aligned}$$

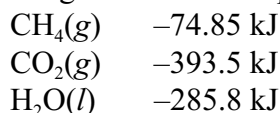
Note that the value of $\Delta H_f^\circ(\text{O}_2)$ is taken as 0, because $\text{O}_2(g)$ is the standard state for elemental oxygen.

As expected by the Law of Hess, the answers are virtually the same. In many cases there is a slight difference, due to minor differences in the values from the experimental data.

4. Calculate the heat of combustion of methane, $\text{CH}_4(\text{g})$, defined by the following thermochemical equation:



Given the following standard enthalpies of formation:



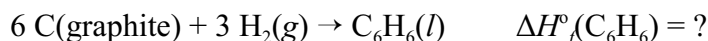
This can be calculated as follows:

$$\begin{aligned}\Delta H^\circ_{\text{rxn}} &= \{\Delta H^\circ_f(\text{CO}_2) + 2\Delta H^\circ_f(\text{H}_2\text{O})\} - \{\Delta H^\circ_f(\text{CH}_4) + 2\Delta H^\circ_f(\text{O}_2)\} \\ &= \{(-393.5 \text{ kJ}) + (2)(-285.8 \text{ kJ})\} - \{(-74.85 \text{ kJ}) + 0\} \\ &= -890.2_5 \text{ kJ} = -890.2 \text{ kJ}\end{aligned}$$

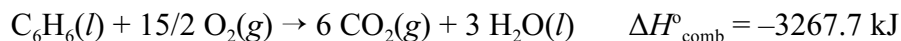
Note that the value of $\Delta H^\circ_f(\text{O}_2)$ is taken as 0, because $\text{O}_2(\text{g})$ is the standard state for elemental oxygen.

5. Calculate the enthalpy of formation, ΔH°_f , for benzene, $\text{C}_6\text{H}_6(\text{l})$, given that the heats of formation of $\text{CO}_2(\text{g})$ and $\text{H}_2\text{O}(\text{l})$ are -393.5 kJ and -285.8 kJ , respectively, and that the heat of combustion of $\text{C}_6\text{H}_6(\text{l})$ is -3267.7 kJ . To do this, carry out the following steps.

- i. Write the balanced thermochemical equation that defines the enthalpy of formation of benzene, $\text{C}_6\text{H}_6(\text{l})$.



- ii. Write the balanced thermochemical equation for the heat of combustion of benzene, $\text{C}_6\text{H}_6(\text{l})$.



- iii. Based on your answer to question ii, write an expression for the heat of combustion of benzene, $\Delta H^\circ_{\text{comb}}$, in terms of the enthalpies of formation of the reactants and products. Using the data given in the problem, solve this for the unknown value of $\Delta H^\circ_f(\text{C}_6\text{H}_6)$, the enthalpy of formation of benzene.

We can express the heat of combustion as

$$\Delta H^\circ_{\text{comb}} = 6\Delta H^\circ_f(\text{CO}_2) + 3\Delta H^\circ_f(\text{H}_2\text{O}) - \Delta H^\circ_f(\text{C}_6\text{H}_6)$$

We have values for everything in this equation except $\Delta H_f^\circ(\text{C}_6\text{H}_6)$. Therefore, we can rearrange and solve for it:

$$\begin{aligned}\Delta H_f^\circ(\text{C}_6\text{H}_6) &= 6\Delta H_f^\circ(\text{CO}_2) + 3\Delta H_f^\circ(\text{H}_2\text{O}) - \Delta H_{\text{comb}}^\circ \\ &= (6)(-393.5 \text{ kJ}) + (3)(-285.8 \text{ kJ}) - (-3267.7 \text{ kJ}) = +49.3 \text{ kJ}\end{aligned}$$

6. For each of the following, indicate which kind of radiation has higher energy.
red light or **blue light**
infrared radiation or radio waves
x-rays or visible light
7. For each of the following, indicate which has higher frequency
light with $\lambda = 490 \text{ nm}$ or light with $\lambda = 520 \text{ nm}$ ($\nu = c/\lambda$)
light with energy of $3.0 \times 10^{-19} \text{ J}$ or light with energy of **$4.5 \times 10^{-19} \text{ J}$** ($E = h\nu$)
8. An argon laser emits green light with a wavelength of 514.5 nm . Calculate the following for this light: (a) the wavelength in \AA ; (b) the frequency in Hz (s^{-1}); (c) the energy in joules, J.

$$(a) \lambda(\text{\AA}) = (514.5 \text{ nm}) \left(\frac{10^{-9} \text{ m}}{\text{nm}} \right) \left(\frac{\text{\AA}}{10^{-10} \text{ m}} \right) = 5145 \text{\AA}$$

Note that to convert from nanometers to \AA ngstroms we need only move the decimal point to the right one place.

- (b) Use $\nu = c/\lambda$. We need wavelength in meters, because the units on the speed of light are $\text{m}\cdot\text{s}^{-1}$. We could convert nm to m first, but with a calculator it is simpler to enter nanometers as 10^{-9} m ; i.e., “514.5E-9”.

$$\nu = \frac{2.9979 \times 10^8 \text{ m}\cdot\text{s}^{-1}}{514.5 \times 10^{-9} \text{ m}} = 5.827 \times 10^{14} \text{ s}^{-1} = 5.827 \times 10^{14} \text{ Hz}$$

- (c) Given the preceding results, we can calculate energy by either $E = h\nu$ or $E = hc/\lambda$. Using the first of these,

$$E = h\nu = (6.626 \times 10^{-34} \text{ J}\cdot\text{s})(5.827 \times 10^{14} \text{ s}^{-1}) = 3.861 \times 10^{-19} \text{ J}$$

9. Is energy emitted or absorbed when the following transitions occur in hydrogen:
- (a) from $n = 1$ to $n = 3$
The transition is to a higher energy state, so energy corresponding to the difference between the states must be absorbed.
- (b) from $n = 5$ to $n = 2$
The transition is to a lower energy state, so energy corresponding to the difference between the states must be emitted.

(c) an H^+ ion acquires an electron into the $n = 2$ state.

An H^+ ion is essentially a hydrogen atom with an electron an infinite distance away, the state for which $n = \infty$. By adding an electron to the $n = 2$ state, it has lower energy (the electron is closer to the nucleus to which it is attracted), so energy will be emitted

10. The four visible lines of the Balmer series in the emission spectrum of hydrogen are violet (410 nm), blue (434 nm), blue-green (486 nm), and red (656 nm). Assign these as state-to-state transitions of the type $n_i \rightarrow n_f$, giving the n values involved for each line.

All transitions in the Balmer series are to $n_f = 2$ from $n_i > 2$. The energy of light increases going from red to violet, so the red line is the smallest state-to-state separation, and the violet line is the largest.

red (656 nm)	$3 \rightarrow 2$
blue-green (486 nm)	$4 \rightarrow 2$
blue (434 nm)	$5 \rightarrow 2$
violet (410 nm)	$6 \rightarrow 2$

11. Calculate the energy of the first line in the Lyman series for the hydrogen atom, which arises from a transition from $n_i = 2$ to $n_f = 1$. What is the wavelength of the radiation emitted? In what region of the electromagnetic spectrum does it fall?

$$\begin{aligned}\Delta E &= (-2.18 \times 10^{-18} \text{ J}) \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right) = (-2.18 \times 10^{-18} \text{ J}) \left(\frac{1}{1^2} - \frac{1}{2^2} \right) \\ &= (-2.18 \times 10^{-18} \text{ J}) \left(\frac{3}{4} \right) = -1.63_5 \times 10^{-18} \text{ J}\end{aligned}$$

$$E_{\text{photon}} = 1.63_5 \times 10^{-18} \text{ J}$$

$$\lambda = \frac{hc}{E} = \frac{(6.626 \times 10^{-34} \text{ J}\cdot\text{s})(2.9979 \times 10^8 \text{ m}\cdot\text{s}^{-1})}{1.63_5 \times 10^{-18} \text{ J}} = 1.21_{49} \times 10^{-7} \text{ m} = 121 \text{ nm}$$

A wavelength of 121 nm falls well within the ultraviolet region.