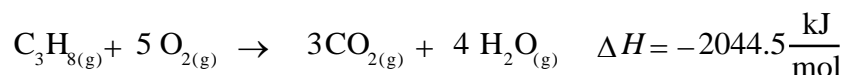


Pretest 2.1 2015 version

Section I Getting ΔH

1. Some automobiles and buses are equipped to burn propane gas, C_3H_8 , as a fuel.

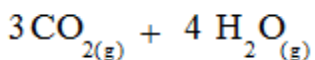
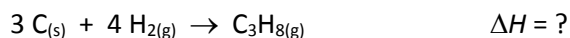
The complete combustion of propane is shown by the following chemical equation:



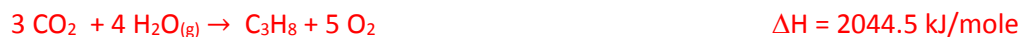
Given the following heats of formation.



a) What is the heat of formation of propane?



Reverse equation (1):



Multiply equation (2) by 4:



Multiple equation (3) by 3:



Add them up:



b) Express your answer in (a) in kJ/mole of C

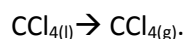
$$\Delta H = -104 \text{ kJ}/3\text{mole C} = -34.7 \text{ kJ/mol C}$$

- c) How much heat would be either released or absorbed (specify) if 4.4 g of propane were formed from carbon and hydrogen?

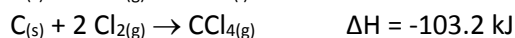
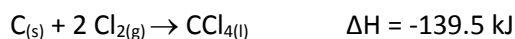
4.4g / (44 g/mole) = 0.10 mole $C_3H_8(g)$
 0.10 mole $C_3H_8(g)$ (-104 kJ/mole) = -10 kJ (released)

2. There is a ΔH (heat of vaporization) associated with the evaporation of $CCl_4(l)$.

- a) Write a simple equation to represent the evaporation of $CCl_4(l)$.



- b) Use the following thermochemical equations to calculate the heat of vaporization of $CCl_4(g)$:



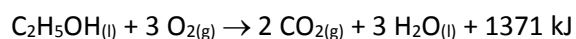
Reverse 1st equation and add it to the second one:

$\Delta H = 36.3 \text{ kJ}$ (notice that the sig fig rule for addition applies; there was no multiplication or division, so we preserve decimal places.)

- c) Plot the enthalpy versus progress of reaction graph and locate $CCl_4(l)$ and $CCl_4(g)$ on the graph. Clearly label both axes.

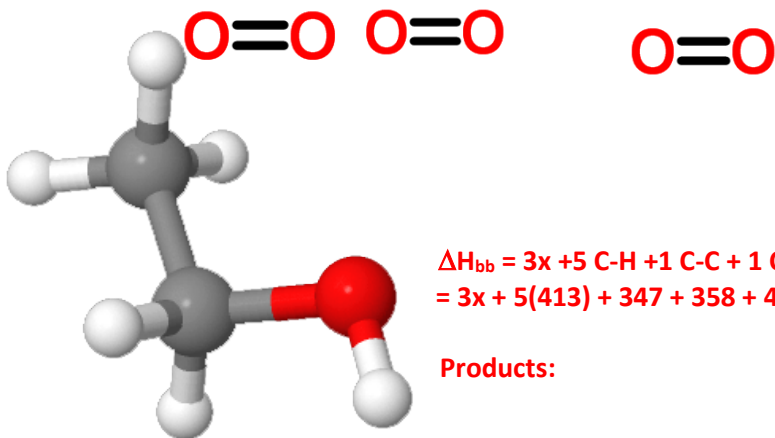
Draw it like 2(b) with $CCl_4(l)$ as a reactant at the bottom of the hill.

3. Consider the following combustion reaction of ethanol:

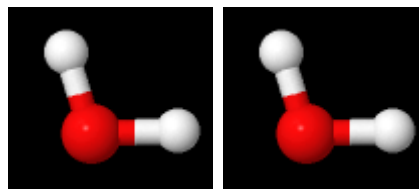
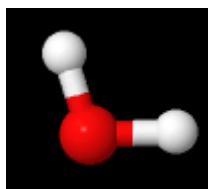


Bond	Energy (kJ/mol)	Bond	Energy (kJ/mol)
H – H	436	C = O	805
H – O	460	C – C	347
H – C	413	C = C	607
C – O	358	C \equiv C	839
O – O	204		

Calculate the energy of the double bond between oxygen (O) atoms in the diatomic oxygen gas (O_2) molecule.



Products:



$$\Delta H_{bf} = 4(\text{C=O}) + 6(\text{O-H}) = 4(-805) + 6(-460) = -5980 \text{ kJ}$$

$$\Delta H_{bb} + \Delta H_{bf} = \Delta H$$

$$3x + 3230 + -5980 = -1371 \text{ kJ (see original equation)}$$

$X = 460. \text{ kJ per } \text{O}_2 \text{ bond (an approx. because throughout intermolecular bonds are ignored and because we assume that all bonds of one type are identical)}$

Section II: Routine Calorimetry Problems(One of each type)

- You only have to burn 1.000 grams of octane(C_8H_{18} -found in gasoline) to raise the temperature of 1000.0 grams of water from 20.00 to 31.42 °C. Calculate the molar heat of combustion of octane.

Environment:

$$\begin{aligned}
 Q &= mc\Delta T \\
 &= 1000 \text{ g}(4.19 \text{ J/g}^\circ\text{C})(31.42 - 20.00)^\circ\text{C} \\
 &= 47849.8000 \text{ J}
 \end{aligned}$$

Reactant:

$$\begin{aligned}
 \Delta H &= -Q = -47.8498 \text{ kJ} \\
 n &= 1.000\text{g}/(114 \text{ g/mole}) = 0.008771929825 \text{ moles} \\
 \Delta H/n &= -47.8498 \text{ kJ}/0.008771929825 \text{ moles} = -5455 \text{ kJ/mole}
 \end{aligned}$$



2. Nacho chips are very calorie-intensive: their ΔH is -20.49 kJ per gram. If only 6.00 grams of chips are burned in a calorimeter containing 5.00 L originally at 19.30 °C, what maximum temperature will the water attain?

Reactant

$$\Delta H = -20.49 \text{ kJ/gram} (6.00 \text{ grams}) = -122.9400 \text{ kJ}$$

Environment:

$$Q = -\Delta H = 122.9400 \text{ kJ} = 122\,940 \text{ J}$$

$$Q = mc\Delta T$$

$$122\,940 \text{ J} = 5000 \text{ ml}(1.0\text{g/ml})(4.19 \text{ J/g}^\circ\text{C})(x - 19.30)^\circ\text{C}$$

$$x = 25.17^\circ\text{C}$$

3. When a calorimeter was filled with 20.0 mL of 3.00 mol/L hydrochloric acid, $\text{HCl}_{(aq)}$, and 50.0 mL of 1.20 mol/L sodium hydroxide, $\text{NaOH}_{(aq)}$, the temperature rose from 22.4°C to 29.8°C.

What was the molar heat of neutralization of $\text{HCl}_{(aq)}$?

(Assume the density and specific heat for all solutions to be equal to that of water, and ignore the water created by the reaction.)

Environment:

$$Q = mc\Delta T$$

$$= (20.0 + 50.0)\text{ml} (1.0 \text{ g/ml}) (4.19 \text{ J/g}^\circ\text{C})(29.8 - 22.4)^\circ\text{C}$$

$$= 2170.4200 \text{ J}$$

Reactant :

$$\Delta H = -Q = -2170.4200 \text{ J} = -2.17042 \text{ kJ}$$

$$n \text{ for HCl} = CV = 3.0 \text{ moles/L} (0.020 \text{ L}) = 0.060 \text{ moles}$$

$$\Delta H/n = -2.17042 \text{ kJ}/0.060 \text{ moles}$$

$$= -36.2 \text{ kJ/mole}$$

- 4.

Specific Heat Capacities of Some Common Substances ($\text{J/g}^\circ\text{C}$)					
Silver	0.238	Lead	0.159	Ice (solid)	2.06
Iron	0.453	Nickel	0.105	Glass	0.84
Copper	0.385	Gold	0.130		

Someone had to find the identity of an unknown solid. The investigator placed this solid in 100.0°C water and allowed it to heat up to the temperature of the water. She then placed the hot metal in a calorimeter containing cool water and collected the data recorded in the data table below.

Mass of the unknown solid	52.8 g
Temperature of the hot water bath	100.0°C
Initial temperature of the water in the calorimeter	21.3°C
Final temperature of the water in the calorimeter	26.0°C
Volume of the water in the calorimeter	90.0 mL

Using the data given and the table of Specific Heat Capacities, determine the identity of the unknown solid. (modified from June 1993)

$$-Q_{\text{hot}} = Q_{\text{cold}}$$

$$-(mc\Delta T)_{\text{hot}} = (mc\Delta T)_{\text{cold}}$$

$$-52.8c(26 - 100) = 90(4.19)(26-21.3)$$

$$c = 0.454 \text{ J/g } ^\circ\text{C}$$

the unknown is iron.

Section III Problems With a Twist or Two

5. Pure 20.5°C alcohol (density 0.791 g/ml; specific heat 2.4 J/g°C) is added to 350 ml of 5.0 °C water to create a mixture with a temperature at 18.0°C. What will be the % of alcohol (by volume) in such a mixture?

$$-Q_{\text{hot}} = Q_{\text{cold}}$$

$$-(mc\Delta T)_{\text{hot}} = (mc\Delta T)_{\text{cold}}$$

$$-x (2.4 \text{ J/g}^\circ\text{C})(18 - 20.5) = 350 \text{ ml (1.0g/ml) (4.19 J/g}^\circ\text{C})(18 - 5) ^\circ\text{C}$$

$$x = 3177 \text{ g}$$

$$= 3177 \text{ g (0.791 g/ml) = 4017 ml}$$

$$\% \text{ alcohol} = 4017 / (4017 + 350) * 100 \% = 92\%$$

6. 500.0 g of 1010 °C copper and 400.0 g of 935°C iron were added to a 2.0 L bucket of ice cold water(0.0°C). How hot did the water get?

Copper's specific heat = 0.39 J/g°C.

Iron's specific heat = 0.45 J/g°C.

$$-Q_{\text{hot 1}} + -Q_{\text{hot 2}} = Q_{\text{cold}}$$

$$-(mc\Delta T)_{\text{hot1}} - (mc\Delta T)_{\text{hot2}} = (mc\Delta T)_{\text{cold}}$$

$$-500 * 0.39 * (x - 1010) - 400 * 0.45(x - 935) = 2000 * 4.19 * (x - 0)$$

$$x = 41.7 ^\circ\text{C}$$

For flashback topics, see strike package.