Solutions to Enthalpy and Thermochemistry Problem Set

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In the video clip from George Gobel you saw that it is possible to light a charcoal grill in about a minute. If the grill is prepared using one 20 pound bag of charcoal (primarily carbon) and two 1 gallon buckets of liquid oxygen (the density o LOX (liquid oxygen) is 1142 kg/m3, CRC Handbook of Chemistry and Physics.)

1.Write a balanced chemical equation describing this reaction.

$C(s) + O_2\,(l) \;\text{-->}\; CO_2\,(g)$

2.Determine the limiting reagent for this reaction.

amount of carbon

$mass_C := 20 \cdot lb$ $\quad mass_C = 9.07185 \cdot 10^3 \cdot gm$

$MW_C := 12.0107 \cdot gm \cdot mole^{-1}$

$mole_C := \frac{mass_C}{MW_C}$ $\quad mole_C = 755.3138 \cdot mole$

amount of oxygen

$V_{O2} := 2 \cdot gal$ $\quad V_{O2} = 7.57082 \cdot 10^3 \cdot mL$

$density_{O2} := 1142 \cdot kg \cdot m^{-3}$ $\quad density_{O2} = 1.142 \cdot gm \cdot mL^{-1}$

$mass_{O2} := V_{O2} \cdot density_{O2}$ $\quad mass_{O2} = 8.64588 \cdot 10^3 \cdot gm$

$MW_{O2} := 2 \cdot 15.9994 \cdot gm \cdot mole^{-1}$ $\quad MW_{O2} = 31.9988 \cdot gm \cdot mole^{-1}$

$mole_{O2} := \frac{mass_{O2}}{MW_{O2}}$ $\quad mole_{O2} = 270.1939 \cdot mole$

Since the equation is balanced with a 1:1 stoichiometry for C and O2, an equal number of moles are required. Si there are fewer moles of O2, this is the limiting reagent.

3.Calculate the mass of each product and any remaining reactant.

$mole_{O2\_final} := mole_{O2} - mole_{O2}$ $\quad mole_{O2\_final} = 0 \cdot mole$

$mass_{O2\_final} := mole_{O2\_final} \cdot MW_{O2}$ $\quad mass_{O2\_final} = 0 \cdot gm$

$mole_{C\_final} := mole_{C} - mole_{O2}$ $\quad mole_{C\_final} = 485.11989 \cdot mole$

$mass_{C\_final} := mole_{C\_final} \cdot MW_{C}$ $\quad mass_{C\_final} = 5.82663 \cdot 10^{3} \cdot gm$

$mole_{CO2\_final} := mole_{O2}$ $\quad mole_{CO2\_final} = 270.1939$

$MW_{CO2} := (12.0107 + 2 \cdot 15.9994) \cdot gm \cdot mole^{-1}$ $\quad MW_{CO2} = 44.0095 \cdot gm \cdot mole^{-1}$

$mass_{CO2} := mole_{O2} \cdot MW_{CO2}$ $\quad mass_{CO2} = 1.18911 \cdot 10^{4} \cdot gm$

4.Calculate the amount of heat released by this reaction.

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

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

$\Delta H_{f\_O2} := -4.812 \cdot kJ \cdot mole^{-1}$ Value for O2 (l) at -80K, not 0 kJ mole-1, since not at standard conditions.

$\Delta H_{f\_CO2} := -393.51 \cdot kJ \cdot mole^{-1}$

$\Delta H_{rxn} := \Delta H_{f\_CO2} \cdot 1 \cdot mole - \left(\Delta H_{f\_C} \cdot 1 \cdot mole + \Delta H_{f\_O2} \cdot 1 \cdot mole\right)$

$\Delta H_{rxn} = -388.698 \cdot kJ$

$mole_{rxn} := mole_{O2}$

$\Delta H := \Delta H_{rxn} \cdot mole_{rxn}$

$\Delta H = -1.05024 \cdot 10^{5} \cdot kJ$

5.If this heat is used to warm a 500 gallon tank of water at 15 C, what is the final temperature?

$E := -\Delta H$ $\quad E = 1.05024 \cdot 10^{5}\ \cdot kJ$

$V_{H2O} := 500 \cdot gal$ $\quad V_{H2O} = 1.89271 \cdot 10^{6}\ \cdot mL$

$density_{H2O} := 999.1026 \cdot kg \cdot m^{-3}$ $\quad density_{H2O} = 0.9991 \cdot gm \cdot mL^{-1}$

$mass_{H2O} := V_{H2O} \cdot density_{H2O}$ $\quad mass_{H2O} = 1.89101 \cdot 10^{6}\ \cdot gm$

$C_{H2O\_l} := 4.2 \cdot joule \cdot gm^{-1} \cdot K^{-1}$

$E = mass \cdot C \cdot \Delta T$

$\Delta T_{H2O} := \frac{E}{mass_{H2O} \cdot C_{H2O\_l}}$ $\quad \Delta T_{H2O} = 13.22347 \cdot K$

$\Delta T = T_{final} - T_{initial}$

$T_{initial\_H2O} := (273.15 + 15) \cdot K$

$T_{final\_H2O} := \Delta T_{H2O} + T_{initial\_H2O}$ $\quad T_{final\_H2O} = 301.37347 \cdot K$

$T_{final\_H2O} \cdot K^{-1} - 273.15 = 28.22347$

6.If this heat is used to warm 40 kg of ice at 200 K, what is the final temperature?

Constants

$C_{H2O\_s} := 2.1 \cdot joule \cdot gm^{-1} \cdot K^{-1}$

$C_{H2O\_g} := 2.0 \cdot joule \cdot gm^{-1} \cdot K^{-1}$

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

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

First step, warm to 273.15 K

$mass_{H2O} := 40 \cdot kg$ $\quad mass_{H2O} = 4 \cdot 10^{4}\ \cdot gm$

$E_{step\_1} := mass_{H2O} \cdot (273.15 \cdot K - 200 \cdot K) \cdot C_{H2O\_s}$

$E_{step\_1} = 6.1446 \cdot 10^{3}\ \cdot kJ$

$E_{remain} := E - E_{step\_1}$ $\quad E_{remain} = 9.88792 \cdot 10^{4}\ \cdot kJ$

Second step, melt ice

$$E_{step\_2} := \Delta H_{fusion} \cdot mass_{H2O} \qquad E_{step\_2} = 1.332 \cdot 10^{4} \cdot kJ$$

$$E_{remain} := E - E_{step\_1} - E_{step\_2} \qquad E_{remain} = 8.55592 \cdot 10^{4} \cdot kJ$$

Third step, heat liquid water

$$E_{step\_3} := mass_{H2O} \cdot (373.15 \cdot K - 273.15 \cdot K) \cdot C_{H2O\_l}$$

$$E_{step\_3} = 1.68 \cdot 10^{4} \cdot kJ$$

$$E_{remain} := E - E_{step\_1} - E_{step\_2} - E_{step\_3} \qquad E_{remain} = 6.87592 \cdot 10^{4} \cdot kJ$$

Fourth step, boil water

$$E_{step\_4} := \Delta H_{vap} \cdot mass_{H2O}$$

$$E_{step\_4} = 9.04 \cdot 10^{4} \cdot kJ$$

$$E_{remain} := E - E_{step\_1} - E_{step\_2} - E_{step\_3} - E_{step\_4} \qquad E_{remain} = -2.16408 \cdot 10^{4} \cdot kJ$$

There is not sufficent energy available to boil all of the water, so only some of the water is converted into stean the final temperature would be 100 C, or 373.15 K.

The mass of water that does boil is determined by the amount of energy available.

$$mass_{H2O\_boil} := \frac{E - E_{step\_1} - E_{step\_2} - E_{step\_3}}{\Delta H_{vap}}$$

$$mass_{H2O\_boil} = 30.42444 \cdot kg$$