

Enthalpy

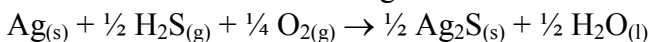
Sample Problems:

1. Given the equation $3 \text{O}_{2(g)} \rightarrow 2 \text{O}_{3(g)}$ $\Delta H = +285.4 \text{ kJ}$, calculate ΔH for the following reaction. $3/2 \text{O}_{2(g)} \rightarrow \text{O}_{3(g)}$.

Answer

Since $3/2 \text{O}_{2(g)} \rightarrow \text{O}_{3(g)}$ is $1/2$ of $3 \text{O}_{2(g)} \rightarrow 2 \text{O}_{3(g)}$ the enthalpy of the reaction will be $1/2$ as well: $1/2 (+285.4 \text{ kJ}) = +142.7 \text{ kJ}$

2. Given the equation: $2\text{Ag}_2\text{S}_{(s)} + 2\text{H}_2\text{O}_{(l)} \rightarrow 4\text{Ag}_{(s)} + 2\text{H}_2\text{S}_{(g)} + \text{O}_{2(g)}$ $\Delta H = +595.5 \text{ kJ}$, calculate ΔH for the following reaction.



Answer

In this problem, the reaction equation has been reversed and divided by four. The new enthalpy value will be the opposite sign and $1/4$ its original value: $-1/4(+595.5 \text{ kJ})$

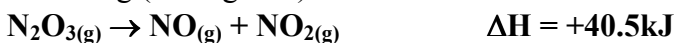
= -148.9 kJ

3. Express the following information as a chemical equation. At 25°C and at a constant pressure, dinitrogen trioxide gas decomposes to nitrogen monoxide and nitrogen dioxide gases with the absorption of 0.533 kJ of heat for every gram of dinitrogen trioxide that decomposes.

Answer

$$+0.533\text{kJ/g N}_2\text{O}_3$$

$$0.533\text{kJ/g (76.01g/mol)} = +40.5\text{kJ/mol}$$



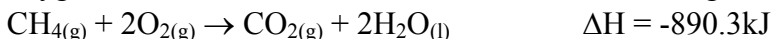
4. What is the enthalpy change when 12.8g $\text{H}_2(\text{g})$ reacts with excess $\text{Cl}_2(\text{g})$ to form $\text{HCl}(\text{g})$?

**Answer**

$$12.8\text{g H}_2 / 2.016\text{g/mol} = 6.35\text{mol H}_2$$

$$6.35\text{mol H}_2 (-184.6\text{kJ/mol H}_2) = -1.17 \times 10^3 \text{kJ}$$

5. What volume of $\text{CH}_4(\text{g})$, measured at 25°C and 745Torr, must be burned in excess oxygen to release $1.00 \times 10^6 \text{kJ}$ of heat to the surroundings?

**Answer**

There is 890.3kJ given off for every mole of CH_4 burned, therefore

the moles of CH_4 needed are $1.00 \times 10^6 \text{kJ} / (890.3\text{kJ/mol}) = 1123\text{mol CH}_4$

$$V = nRT/P = (1123\text{mol})(0.0821)(25^\circ\text{C} + 273.15) / (745/760) = 2.80 \times 10^4 \text{L CH}_4$$

Calorimetry

1. Calculate the heat capacity of a sample of brake fluid if the sample must absorb 911J of heat for its temperature to rise from 15°C to 100°C .

Answer

$$C = q/\Delta T = 911 / (100 - 15) = 10.7\text{J}/^\circ\text{C}$$

2. A burner on an electric range has a heat capacity of 345J/K. What is the value of q , in kilojoules, as the burner cools from a temperature of 467°C to a room temperature of 23°C?

Answer

$$q = C\Delta T = (345\text{J/K})(23-467) = -1.53 \times 10^5 \text{J} = \mathbf{-153\text{kJ}}$$

Molar Heat Capacity/Specific Heat

1. How much heat, in calories and kilocalories, does it take to raise the temperature of 814g of water from 18.0°C to 100°C?

Answer

$$q = mc\Delta T = (814\text{g})(1\text{cal/g}^\circ\text{C})(100^\circ\text{C} - 18^\circ\text{C}) = \mathbf{6.67 \times 10^4 \text{cal} = 66.7\text{Cal (kcal)}}$$

2. What mass of water, in kilograms, can be heated from 5.5°C to 55.0°C by 9.09×10^{10} J of heat?

Answer

$$m = q/c\Delta T = 9.09 \times 10^{10} \text{J} / (4.180\text{J/g}^\circ\text{C})(55.0^\circ\text{C} - 5.5^\circ\text{C}) = 4.39 \times 10^8 \text{g} = \mathbf{4.39 \times 10^5 \text{kg}}$$

3. A 454g block of lead is at an initial temperature of 22.5°C. What will be the temperature of the lead after it absorbs 4.22kJ of heat from its surroundings?

Answer

$$\Delta T = q / mc = 4.22 \times 10^3 \text{J} / (454\text{g})(0.128\text{J/g}^\circ\text{C}) = 72.6^\circ\text{C}$$

Since heat is absorbed the temperature will go up

$$\text{The final temperature is } 22.5^\circ\text{C} + 72.6^\circ\text{C} = \mathbf{95.1^\circ\text{C}}$$

4. How many grams of copper can be heated from 22.5°C to 35.0°C by the same quantity of heat that can raise the temperature of 145g of H₂O from 22.5°C to 35.0°C?

Answer

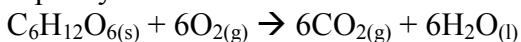
$$(145\text{g})(4.180\text{J/g}^\circ\text{C}) = m(0.385\text{J/g}^\circ\text{C})$$

$$1574.3\text{g}$$

$$\mathbf{1.57\text{kg}}$$

Bomb Calorimetry

1. A 0.8082g sample of glucose (C₆H₁₂O₆) is burned in a bomb calorimeter assembly, and the temperature is noted to rise from 25.11°C to 27.21°C. Determine the heat capacity of the bomb calorimeter assembly, given:



$$\Delta H = -2803\text{kJ}$$

Answer

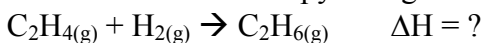
$$\text{moles of glucose} = 0.8082\text{g} / 180.2\text{g/mol} = 0.00449\text{mol}$$

$$q = -2803\text{kJ/mol} (0.00449\text{mol}) = -12.59\text{kJ}$$

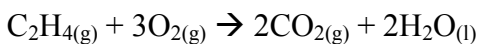
$$C = q/\Delta T = 12.59\text{kJ}/2.10^\circ\text{C} = \mathbf{6.00\text{kJ}/^\circ\text{C}}$$

Hess's Law of Heat Summation

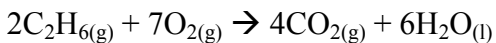
1. Calculate the enthalpy change for the reaction



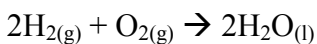
Given:



$$\Delta H = -1410.9\text{kJ}$$

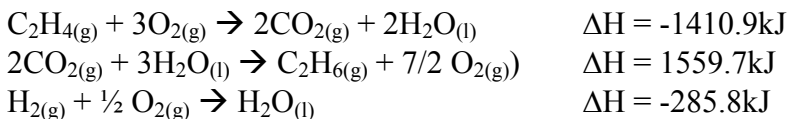
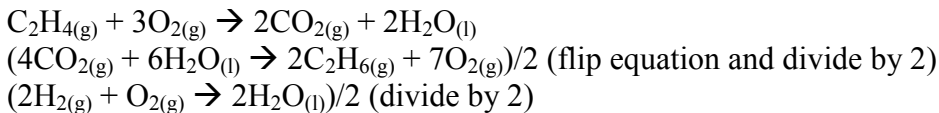


$$\Delta H = -3119.4\text{kJ}$$



$$\Delta H = -571.6\text{kJ}$$

Answer:



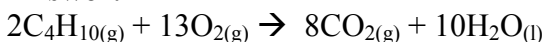
Resulting Equation:



Standard Enthalpy of Formation

1. Use data from table 6.2 (p257) to calculate ΔH° for the combustion of butane gas, C_4H_{10} , to produce gaseous carbon dioxide and liquid water.

Answer:



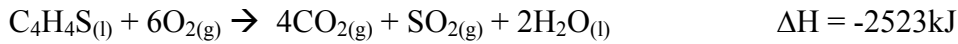
Using

$$\Delta\text{H}^\circ = \sum v_p \times \Delta\text{H}_f^\circ(\text{products}) - \sum v_r \times \Delta\text{H}_f^\circ(\text{reactants})$$

$$\begin{aligned} [10(-285.8\text{kJ/mol}) + 8(-393.5\text{kJ/mol})] - [2(-125.7\text{kJ/mol}) + 13(0\text{kJ/mol})] \\ -6006.0 - (-251.4) = \mathbf{-5754.6\text{kJ}} \end{aligned}$$

2. The combustion of thiophene, $\text{C}_4\text{H}_4\text{S}(\text{l})$, a compound used in the manufacture of pharmaceuticals, produces carbon dioxide and sulfur dioxide gases and liquid water. The enthalpy change in the combustion of one mole of $\text{C}_4\text{H}_4\text{S}(\text{l})$ is -2523kJ . Use this information and data from Table 6.2 to establish ΔH_f° for $\text{C}_4\text{H}_4\text{S}(\text{l})$.

Answer:



$$[4(-393.5) + 1(-296.8) + 2(-285.8)] - [1(\Delta\text{H}) + 6(0)] = -2523$$

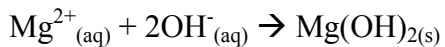
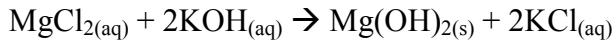
$$-2442.4 - \Delta\text{H} = -2523$$

$$\Delta\text{H} = 80.6\text{kJ/mol}$$

Ionic Reactions in Solution

Given that $\Delta\text{H}^\circ_f[\text{Mg}(\text{OH})_{2(s)}] = -924.5\text{kJ/mol}$, what is the standard enthalpy change, ΔH° , for the reaction of aqueous solutions of magnesium chloride and potassium hydroxide? (Use table 6.3, p 261)

Answer:



$$[1(-924.5\text{kJ})] - [1(-466.9\text{kJ}) + 2(-230.0\text{kJ})]$$

$$\Delta\text{H}^\circ = 2.4\text{kJ per mole of Mg}(\text{OH})_{2(s)} \text{ formed}$$