Enthalpy Stoichiometry

Chem Worksheet 16-3

Name KEY

The **molar enthalpy of reaction** (ΔH_{rxn}) is the amount of heat transferred during a reaction. It is reported in kilojoules per mole of reactant. A reaction that produces heat is **exothermic** and has a negative ΔH_{rxn}. A reaction that absorbs heat is **endothermic** and has a positive ΔH_{rxn}.

Example

How much heat is produced when 85 g of sulfur reacts according to the reaction below?

$$2S + 3O_2 \rightarrow 2SO_3 \quad \Delta H = -792 \text{ kJ}$$

- the ΔH value given in the equation is the amount of heat transferred when **2 moles** of sulfur and **3 moles** of oxygen react.

- write the 'given' and 'unknown' units: $\frac{85\text{ g S}}{1} \times \frac{\quad}{\quad} \times \frac{\quad}{\quad} = \quad \text{kJ}$

- fill in factors: $\frac{85\text{ g S}}{1} \times \frac{1\text{ mol S}}{32.06\text{ g S}} \times \frac{-792\text{ kJ}}{2\text{ mol S}} = \quad \text{kJ}$

- solve: $\frac{85\text{ g S}}{1} \times \frac{1\text{ mol S}}{32.06\text{ g S}} \times \frac{-792\text{ kJ}}{2\text{ mol S}} = -1050 \text{ kJ}$

Answer the following questions. Show all work and report answers with units.

1. How much heat will be released when 6.44 g of sulfur reacts with excess O_2 according to the following equation?

$$2\text{ S} + 3O_2 \rightarrow 2SO_3 \quad \Delta H = -791.4\text{ kJ}$$

$6.44\text{g S} \cdot \frac{1\text{ mol S}}{32.1\text{g S}} \cdot \frac{-791.4\text{ KJ}}{2\text{ mol S}} = \boxed{-79.4\text{ KJ}}$

2. How much heat will be released when 4.72 g of carbon reacts ~~with excess O_2~~ according to the following equation?

$$C + O_2 \rightarrow CO_2 \quad \Delta H = -393.5\text{ kJ}$$

$4.72\text{g C} \left| \frac{1\text{ mol C}}{12\text{g C}} \right| \frac{-393.5\text{ KJ}}{1\text{ mol C}} = \boxed{-154.8\text{ KJ}}$

3. How much heat will be absorbed when 38.2 g of bromine reacts ~~with excess H_2~~ according to the following equation?

$$H_2 + Br_2 \rightarrow 2HBr \quad \Delta H = +72.80\text{ kJ}$$

$38.2\text{g Br} \left| \frac{1\text{ mol Br}}{160\text{g Br}} \right| \frac{72.80\text{ KJ}}{1\text{ mol Br}_2} = \boxed{17.4\text{ KJ}}$

4. How much heat will be released when 1.48 g of chlorine reacts with excess phosphorus according to the following equation.

$$2P + 5Cl_2 \rightarrow 2PCl_5 \quad \Delta H = -886\text{ kJ}$$

$1.48\text{g Cl}_2 \left| \frac{1\text{ mol Cl}_2}{70\text{g Cl}_2} \right| \frac{-886\text{ KJ}}{5\text{ mol Cl}_2} = \boxed{-3.75\text{ KJ}}$

5. What mass of propane, C_3H_8 must be burned in order to produce 76,000 kJ of energy?

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O \quad \Delta H = -2200\text{ kJ}$$

$-76000\text{ KJ} \cdot \frac{1\text{ mol } C_3H_8}{-2200\text{ KJ}} \cdot \frac{44\text{g } C_3H_8}{1\text{ mol } C_3H_8} = \boxed{1520\text{ g}}$

6. How much heat will be absorbed when 13.7 g of nitrogen reacts ~~with excess O_2~~ according to the following equation?

$$N_2 + O_2 \rightarrow 2NO \quad \Delta H = +180\text{ kJ}$$

$13.7\text{g } N_2 \left| \frac{1\text{ mol } N_2}{28\text{g } N_2} \right| \frac{180\text{ KJ}}{1\text{ mol } N_2} =$

7. What mass of iron must react to produce 3600 kJ of energy?

$$3Fe + 2O_2 \rightarrow Fe_3O_4 \quad \Delta H = -1120\text{ kJ}$$

$-3600\text{KJ} \left| \frac{3\text{ mol Fe}}{-1120\text{KJ}} \right| \frac{55.8\text{g Fe}}{1\text{ mol Fe}} = \boxed{538.1\text{ g Fe}}$

8. How much heat will be released when 12.0 g of H_2 reacts with 76.0 g of O_2 according to the following equation?
(when one reactant runs out the reaction stops)

$$2H_2 + O_2 \rightarrow 2H_2O \quad \Delta H = -571.6\text{ kJ}$$

$12.0\text{g } H_2 \left| \frac{1\text{ mol } H_2}{2\text{ g } H_2} \right| \frac{-571.6\text{ KJ}}{2\text{ mol } H_2} = \boxed{-1714.8\text{ KJ}}$

O_2 will run out 1st

$76.0\text{g } O_2 \left| \frac{1\text{ mol } O_2}{32.0\text{g } O_2} \right| \frac{-571.6\text{ KJ}}{1\text{ mol } O_2} = \boxed{-1357.6\text{ KJ}}$

Enthalpy Worksheet

1. The combustion of methane, CH_4, releases $890.4\ kJ/mol$ of heat. That is, when one mole of methane is burned, $890.4\ kJ$ are given off to the surroundings. This means that the products have $890.4\ kJ$ less energy stored in the bonds than the reactants. Thus, ΔH for the reaction $= -890.4\ kJ$. A negative symbol for ΔH indicates an exothermic reaction.

$$CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(L)} \qquad \Delta H = -890.4\ kJ$$

A. How much energy is given off when $2.00\ mol$ of CH_4 are burned?

$$\frac{2.00\ mol\ CH_4}{1} \left| \frac{-890.4\ KJ}{1\ mol\ CH_4} \right. = \boxed{-1780.8\ KJ}$$

B. How much energy is released when $22.4\ g$ of CH_4 are burned?

$$\frac{22.4\ g\ CH_4}{16\ g\ CH_4} \left| \frac{1\ mol\ CH_4}{1} \right| \frac{-890.4\ KJ}{1\ mol\ CH_4} = \boxed{-1246.6\ KJ}$$

C. If you were to attempt to make $45.0\ g$ of methane from CO_2 and H_2O (with O_2 also being produced), how much heat would be absorbed during the reaction?

$$\frac{45\ g\ CH_4}{16\ g\ CH_4} \left| \frac{1\ mol\ CH_4}{1} \right| \frac{890.4\ KJ}{1\ mol\ CH_4} = \boxed{2504.2\ KJ}$$