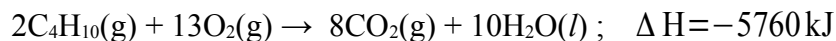


CHM 110 -Heat Practice Set**SOLUTIONS****Solve the problems.**

1) If 1.54 L of butane (C_4H_{10}) at $25^\circ C$ and 1.00 atm is burned, how much heat is evolved?



- 181 kJ heat evolved.

Complete solution:

First, use $PV = nRT$ to find the moles of butane reacted.

$$P = 1.00 \text{ atm}$$

$$n = ?$$

$$V = 1.54 \text{ L}$$

$$T = 25^\circ C = 298 \text{ K}$$

$$n = \frac{(1 \text{ atm}) \times (1.54 \text{ L})}{(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}) \times (298 \text{ K})} = 0.06297569 \text{ mol } C_4H_{10}$$

Next, use the thermochemical equation to find the heat (enthalpy change).

$$0.06297569 \text{ mol } C_4H_{10} \times \frac{-5760 \text{ kJ}}{2 \text{ mol } C_4H_{10}} = -181 \text{ kJ}$$

Note: Since the word “**evolved**” implies the process is exothermic (heat is leaving the system), we say that the “heat evolved” is 181 kJ. The enthalpy change for the process is -181 kJ.

2) If 2.57 g of Na_2O_2 is reacted with water, how much heat is evolved?



- 4.73 kJ heat evolved.

Complete solution:

Find the number of moles of sodium peroxide using the formula weight of sodium peroxide, then find heat using stoichiometry.

$$2.57 \text{ g Na}_2\text{O}_2 \times \frac{1 \text{ mol}}{77.98 \text{ g}} \times \frac{-287 \text{ kJ}}{2 \text{ mol Na}_2\text{O}_2} = -4.73 \text{ kJ}$$

Note: As in problem #1, this one asks us about heat *evolved*, so we report 4.73 kJ as our answer. The enthalpy change for the process is -4.73 kJ.

3) Calculate (from heats of formation) the enthalpy change for the following reaction:



• $\Delta H = \underline{202.4} \text{ kJ}$

Complete solution:

Use Hess's Law to find the enthalpy change. Find standard enthalpies of formation in the back of your textbook or via the Internet. Be careful; make sure the phase labels on your enthalpies of formation match the ones in the chemical equation.

<i>Reactants</i>	<i>Products</i>
$\text{Cl}_2(\text{g}); \Delta H_{\text{f}} = 0 \text{ kJ/mol}$	$\text{HCl}(\text{g}); \Delta H_{\text{f}} = -92.3 \text{ kJ/mol}$
$\text{H}_2\text{O}(\text{l}); \Delta H_{\text{f}} = -285.8 \text{ kJ/mol}$	$\text{O}_2(\text{g}); \Delta H_{\text{f}} = 0 \text{ kJ/mol}$

Reactants: $2 \times (0 \text{ kJ}) + 2 \times (-285.8 \text{ kJ}) = -571.6 \text{ kJ}$

Products: $4 \times (-92.3 \text{ kJ}) + 1 \times (0 \text{ kJ}) = -369.2 \text{ kJ}$

Reaction = Products - Reactants = 202.4 kJ