

CONCEPT: ENTHALPY OF FORMATION

- At this point we have identified three ways to determine the enthalpy or heat of reaction, ΔH_{Rxn} .
 - Constant Volume Calorimetry:** Uses a bomb calorimeter to find ΔH_{Rxn} of _____ reactions.
 - Thermochemical Equation:** Uses _____ and a _____ equation to determine the ΔH_{Rxn} .
 - Hess's Law:** Uses the enthalpies of _____ reactions to find ΔH_{Rxn} for the overall reaction.
- If the first 3 ways are unavailable, then we can use the standard enthalpy of formation for substances to find ΔH_{Rxn} .
 - Recall, that an element in its standard state is given an enthalpy of formation of _____.

Standard Heat of Reaction

Standard Heat of Reaction Formula

$$\Delta H_{\text{Rxn}}^{\circ} = \left[\sum n \Delta H_f^{\circ} (\text{products}) \right] - \left[\sum n \Delta H_f^{\circ} (\text{reactants}) \right]$$

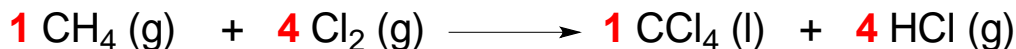
$\Delta H_{\text{Rxn}}^{\circ}$ = Standard enthalpy or heat of _____ in kJ.

\sum = sigma or "sum up".

n = _____ of substance.

ΔH_f° = Standard enthalpy or heat of _____ in $\frac{\text{kJ}}{\text{mol}}$.

EXAMPLE: The reaction of methane with chlorine gas is illustrated by the reaction below:



Calculate the $\Delta H_{\text{rxn}}^{\circ}$ if the standard enthalpies of formation for CH_4 , CCl_4 , and HCl are -74.87 kJ/mol , -139 kJ/mol and -92.31 kJ/mol respectively.

STEP 0: CHECK to see if the chemical equation is balanced and if not then do the necessary steps to balance it.

STEP 1: Starting with the products, multiple the **coefficients** of each substance with their enthalpy of formation value.

$$\text{Products} = \left[\left(\text{--- mol CCl}_4 \times \text{---} \frac{\text{kJ}}{\text{mol}} \right) + \left(\text{--- mol HCl} \times \text{---} \frac{\text{kJ}}{\text{mol}} \right) \right] =$$

STEP 2: Go to the reactants, also multiple the **coefficients** of each substance with their enthalpy of formation value.

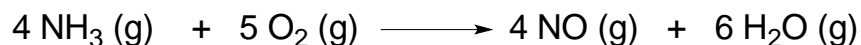
$$\text{Reactants} = \left[\left(\text{--- mol CH}_4 \times \text{---} \frac{\text{kJ}}{\text{mol}} \right) + \left(\text{--- mol Cl}_2 \times \text{---} \frac{\text{kJ}}{\text{mol}} \right) \right] =$$

STEP 3: Take both totals and place them into the standard heat of reaction formula to determine $\Delta H_{\text{Rxn}}^{\circ}$.

$$\Delta H_{\text{Rxn}}^{\circ} = \text{Products} - \text{Reactants} = \left[\text{---} \right] - \left[\text{---} \right] =$$

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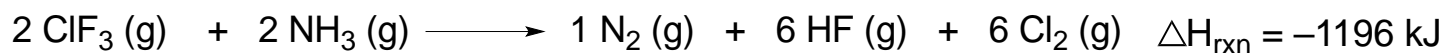
PRACTICE: The oxidation of ammonia is illustrated by the following equation:



Calculate the enthalpy of reaction, ΔH_{Rxn} , based on the given standard heats of formation.

Standard Heats of Formation	
Substances	DH_f° kJ/mol
$\text{NH}_3 (\text{g})$	− 45.9
$\text{O}_2 (\text{g})$	0.0
$\text{NO} (\text{g})$	90.3
$\text{H}_2\text{O} (\text{g})$	− 241.8
$\text{H}_2\text{O} (\text{l})$	− 285.8

PRACTICE: Consider the following equation:



Determine the standard enthalpy of formation for chlorine trifluoride, ClF_3 .

Standard Heats of Formation	
Substances	DH_f° kJ/mol
$\text{NH}_3 (\text{g})$	− 45.9
$\text{N}_2 (\text{g})$	0.0
$\text{HF} (\text{g})$	− 273
$\text{Cl}_2 (\text{g})$	0.0