

## CONCEPT: ENTHALPY OF FORMATION

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

### Standard Heat of Reaction

Standard Heat of Reaction Formula

$$\Delta H_{\text{Rxn}}^{\circ} = [S \mathbf{n} \Delta H_f^{\circ} (\text{_____})] - [S \mathbf{n} \Delta H_f^{\circ} (\text{_____})]$$

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

□  $S$  = sigma or "sum up".

□  $\mathbf{n}$  = \_\_\_\_\_ of substance.

□  $\Delta H_f^{\circ}$  = 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  $\text{CH}_4$ ,  $\text{CCl}_4$ , and  $\text{HCl}$  are  $-74.87 \text{ kJ/mol}$ ,  $-139 \text{ kJ/mol}$  and  $-92.31 \text{ kJ/mol}$  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.

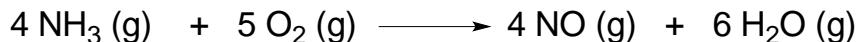
$$\text{Reactants} = \left[ \left( \text{_____ mol CH}_4 \times \text{_____ } \frac{\text{kJ}}{\text{mol}} \right) + \left( \text{_____ mol Cl}_2 \times \text{_____ } \frac{\text{kJ}}{\text{mol}} \right) \right] =$$

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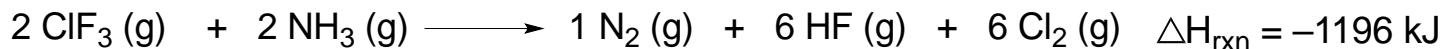
**PRACTICE:** The oxidation of ammonia is illustrated by the following equation:



Calculate the enthalpy of reaction,  $\Delta H_{\text{rxn}}$ , based on the given standard heats of formation.

Standard Heats of Formation	
Substances	$\Delta H_f^\circ$ kJ/mol
$\text{NH}_3 \text{ (g)}$	– 45.9
$\text{O}_2 \text{ (g)}$	0.0
$\text{NO (g)}$	90.3
$\text{H}_2\text{O (g)}$	– 241.8
$\text{H}_2\text{O (l)}$	– 285.8

**PRACTICE:** Consider the following equation:



Determine the standard enthalpy of formation for chlorine trifluoride,  $\text{ClF}_3$ .

Standard Heats of Formation	
Substances	$\Delta H_f^\circ$ kJ/mol
$\text{NH}_3 \text{ (g)}$	– 45.9
$\text{N}_2 \text{ (g)}$	0.0
$\text{HF (g)}$	– 273
$\text{Cl}_2 \text{ (g)}$	0.0