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

$$\Delta H^{o}_{Rxn} = \left[\Sigma n \Delta H^{o}_{f\,(\_\_\_\_\_\_\_\_)}\right] - \left[\Sigma n \Delta H^{o}_{f\,(\_\_\_\_\_\_\_\_)}\right]$$

- □ ΔH^{o}_{Rxn} = Standard enthalpy or heat of __________ in kJ.
- □ Σ = sigma or "sum up".
- □ n = __________ of substance.
- □ ΔH^{o}_{f} = Standard enthalpy or heat of __________ in $\frac{kJ}{mol}$.

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

$$1\ CH_4\ (g) \;+\; 4\ Cl_2\ (g) \longrightarrow 1\ CCl_4\ (l) \;+\; 4\ HCl\ (g)$$

Calculate the ΔH^{o}_{rxn} 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.

$$\text{Products} = \left[\left(\_\_\ \text{mol}\ CCl_4 \times \_\_\_\_\_\_\_\_ \frac{kJ}{mol}\right) + \left(\_\_\ \text{mol}\ HCl \times \_\_\_\_\_\_\_\_ \frac{kJ}{mol}\right)\right] =$$

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

$$\text{Reactants} = \left[\left(\_\_\ \text{mol}\ CH_4 \times \_\_\_\_\_\_\_\_ \frac{kJ}{mol}\right) + \left(\_\_\ \text{mol}\ Cl_2 \times \_\_\_\_\_\_\_\_ \frac{kJ}{mol}\right)\right] =$$

STEP 3: Take both totals and place them into the standard heat of reaction formula to determine ΔH^{o}_{Rxn}.

$$\Delta H^{o}_{Rxn} = \text{Products} - \text{Reactants} = \left[\_\_\_\_\_\_\_\_\right] - \left[\_\_\_\_\_\_\_\_\right] =$$

CONCEPT: ENTHALPY OF FORMATION

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

$$4\ NH_3\ (g) + 5\ O_2\ (g) \longrightarrow 4\ NO\ (g) + 6\ H_2O\ (g)$$

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

Standard Heats of Formation	
Substances	DH_f^o kJ / mol
NH_3 (g)	– 45.9
O_2 (g)	0.0
NO (g)	90.3
H_2O (g)	– 241.8
H_2O (l)	– 285.8

PRACTICE: Consider the following equation:

$$2\ ClF_3\ (g) + 2\ NH_3\ (g) \longrightarrow 1\ N_2\ (g) + 6\ HF\ (g) + 6\ Cl_2\ (g) \quad \triangle H_{rxn} = -1196\ kJ$$

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

Standard Heats of Formation	
Substances	DH_f^o kJ / mol
NH_3 (g)	– 45.9
N_2 (g)	0.0
HF (g)	– 273
Cl_2 (g)	0.0