

Molar Enthalpies of Formation

- Recall, that Hess's law is used to calculate the enthalpy change for a given reaction when a list of intermediate reactions are given
- You can also use **standard molar enthalpies of formation ($\Delta_f H^\circ$)** to calculate the enthalpy of any reaction
 - Every compound should have an enthalpy change for that compound being formed from its elements (ie. a formation reaction).
 - * Therefore, every compound will have a molar enthalpy of formation ($\Delta_f H^\circ$). A list of some compounds and their corresponding molar enthalpies of formation are found on pg. 4-5 of the data book.
 - Molar enthalpies of formation represent the amount of energy required or released when 1 mole of compound is formed from its elements.
 - Example: $\Delta_f H^\circ = -103.8 \text{ kJ/mol}$ for C_3H_8



- We can use these molar enthalpies of formation values to calculate the enthalpy change (ΔH) for any given reaction using the following equation:

not just formation rxns!

$$\Delta H^\circ = \sum_{\text{products}} (n \cdot \Delta_f H^\circ) - \sum_{\text{reactants}} (n \cdot \Delta_f H^\circ)$$

similar to $\Delta H = n \Delta_f H$!

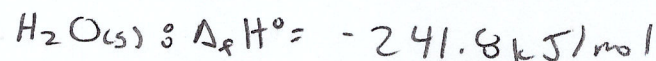
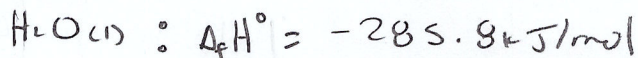
means "sum of" or "total"

where n is the number of moles as found in the balanced reaction (mol)

$\Delta_f H^\circ$ is the molar enthalpy of formation of a compound (kJ/mol)

- Recall the definition of enthalpy change; the difference in potential energy of the final state (products) and the initial state (reactants)
- * When using this formula, the standard molar enthalpies of formation for all elements at SATP have a value of zero because elements can't undergo a formation reaction
 - $\Delta_f H^\circ = 0$ for any element (ie, C, H_2 , Ag)

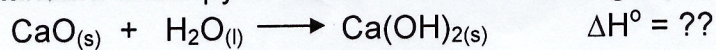
- The phase (solid, liquid, or gas) of the compounds is important because a different phase will change the standard molar enthalpy of formation ($\Delta_f H^\circ$)
 - Example:



- * Calculating the enthalpy change for any given reaction using standard molar enthalpies of formation is useful when you feel like you are missing information and all you have, or can create, is a balanced reaction!

EXAMPLES:

1. Calculate the standard enthalpy of reaction for the following equation



$$\Delta H^\circ = \sum_{\text{prod.}} (n \cdot \Delta_f H^\circ) - \sum_{\text{react.}} (n \cdot \Delta_f H^\circ)$$

$$\Delta H = \left[(n \cdot \Delta_f H^\circ)_{\text{Ca(OH)}_2(s)} \right] - \left[(n \cdot \Delta_f H^\circ)_{\text{CaO}(s)} + (n \cdot \Delta_f H^\circ)_{\text{H}_2\text{O}(l)} \right]$$

↑
↑
 sum of all products sum of all reactants

$$\Delta H = \left[(1 \text{ mol})(-985.2 \text{ kJ/mol}) \right] - \left[(1 \text{ mol})(-634.9 \text{ kJ/mol}) + (1 \text{ mol})(-285.8 \text{ kJ/mol}) \right]$$

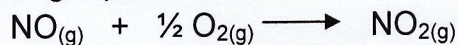
Ca(OH)₂(s)
CaO
H₂O

$$\Delta H = [-985.2 \text{ kJ}] - [-920.7 \text{ kJ}]$$

$$\boxed{\Delta H = -64.5 \text{ kJ}}$$

* moles (n) come from balanced rxn!

2. How much energy is released in the production of 2.37g of nitrogen dioxide according to the following equation?



Specific amount	balanced rxn
$n =$	$n = 1 \text{ mol}$
$\Delta H = ?$	$\Delta H =$ use molar enthalpies of formation

aside

$$n = \frac{m}{M}$$

$$n = \frac{2.37 \text{ g}}{46.01 \text{ g/mol}}$$

$$n = 0.0515 \text{ mol}$$

②

$\Delta_r H$

①

aside

$$\Delta H^\circ = \sum (n \cdot \Delta_f H^\circ)_{\text{prod.}} - \sum (n \cdot \Delta_f H^\circ)_{\text{react.}}$$

$$\Delta H^\circ = [(1 \text{ mol})(+33.2 \text{ kJ/mol})] -$$

NO_2

$$[(1 \text{ mol})(+91.3 \text{ kJ/mol}) + (\frac{1}{2} \text{ mol})(0 \text{ kJ/mol})]$$

NO

O_2

$$\Delta H = [+33.2 \text{ kJ}] - [+91.3 \text{ kJ}]$$

sum of products

→ sum of all reactants

$$\Delta H = -58.1 \text{ kJ}$$

$$\therefore \textcircled{1} \Delta_r H = \frac{\Delta H}{n} = \frac{-58.1 \text{ kJ}}{1 \text{ mol}} = -58.1 \text{ kJ/mol}$$

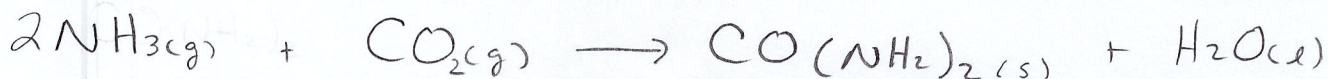
$$\textcircled{2} \Delta H = n \Delta_r H = (0.0515 \text{ mol})(-58.1 \text{ kJ/mol})$$

$$\Delta H = -2.99 \text{ kJ}$$

3. The fertilizer urea ($\text{CO}(\text{NH}_2)_2$) is produced along with liquid water by the reaction of ammonia and carbon dioxide. If the molar enthalpy change for urea ($\text{CO}(\text{NH}_2)_2$) is -134.0 kJ/mol , calculate the standard molar enthalpy of formation for urea ($\text{CO}(\text{NH}_2)_2$)?

$$\hookrightarrow \Delta_f H^\circ = ?$$

* need to work backwards



for the rxn, not the formation rxn. $\text{CO}(\text{NH}_2)_2$ $\rightarrow \Delta_r H = -134.0 \text{ kJ/mol}$

$$\therefore \Delta H = n \Delta_r H = (1.0 \text{ mol})(-134.0 \text{ kJ/mol}) = -134.0 \text{ kJ}$$

for rxn

$$\Delta H = \sum_{\text{prod.}} (n \cdot \Delta_f H^\circ) - \sum_{\text{react.}} (n \cdot \Delta_f H^\circ)$$

$$-134.0 \text{ kJ} = [(1 \text{ mol})(\Delta_f H^\circ) + (1 \text{ mol})(-285.8 \text{ kJ/mol})] -$$

\swarrow sum of products

$$[(2 \text{ mol})(-45.9 \text{ kJ/mol}) + (1 \text{ mol})(-393.5 \text{ kJ/mol})]$$

\nwarrow sum of reactants

$$-134.0 \text{ kJ} = [(1 \text{ mol})(\Delta_f H^\circ) - 285.8 \text{ kJ}] - [-485.3 \text{ kJ}]$$

$$-333.5 \text{ kJ} = (1 \text{ mol})(\Delta_f H^\circ) \quad \begin{matrix} \uparrow \\ \text{products} \end{matrix} \quad \begin{matrix} \uparrow \\ \text{reactants} \end{matrix}$$

$$\boxed{-333.5 \text{ kJ/mol} = \Delta_f H^\circ}$$

Now try pg. 382 # 7-12 & Practice Problems

\downarrow
for #7, simply determine the ΔH°

Practice Problems

1. What is the molar enthalpy change for the combustion of methanol (CH_3OH) in an open system? **[-637.9 kJ/mol]**
2. Given that the molar enthalpy of combustion for propan-1-ol ($\text{C}_3\text{H}_7\text{OH}_{(l)}$) is -2021 kJ/mol in a closed system, what is the molar enthalpy of formation for propan-1-ol? **[-302.7 kJ/mol]**