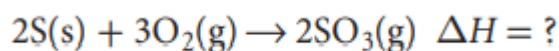


Section 4 : Calculating Enthalpy Change

Hess's Law :

- The change in enthalpy can be measured directly by a calorimeter or indirectly by (theoretical) Hess's law.
- Sometimes it is impossible or impractical to calculate ΔH in a reaction using a calorimeter for the following reasons:
 1. The reaction may occur so slowly that calculating the enthalpy change becomes impossible, such as the reaction of converting carbon from allotropic diamond to allotropic graphite.

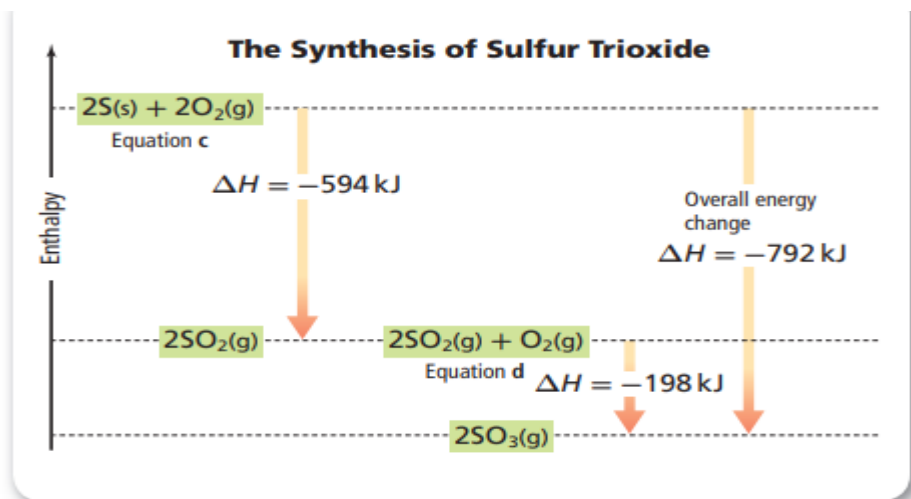
$$\text{C(s, ماس)} \rightarrow \text{C(s, جرافيت)}$$
 2. Other reactions occur under conditions that are difficult to replicate in the laboratory.
 3. There are also other reactions that result in products other than those desired.



- Unfortunately, laboratory experiments to produce sulfur trioxide and determine its ΔH result in a mixture of products that is mostly sulfur dioxide (SO_2).
- That is why chemists use a theoretical method to calculate ΔH , which is Hess's law.
- **Hess's law states :** that if you can add two or more thermochemical equations to produce a final equation for a reaction, then the sum of the enthalpy changes for the individual reactions is the enthalpy change for the final reaction.

➤ **OR**

- When reactants are converted to products, the change in enthalpy is the same whether the reaction takes place in one step or in a series of steps.



➤ Application of Hess' Law:

- 1- We compare the sub-equations with the overall equation so that the reactants and products are identical to the total equation as well as the number of moles.
- 2- If you need to reverse the reaction, the sign of the enthalpy change (ΔH) must be changed.
- 3- When adjusting the number of moles, we multiply the fixed number on all sides of the equation up to (ΔH).
- 4- After adjusting the sub-equations, we sum the equations algebraically.
- 5- When algebraic addition, we cancel the materials in two different sides and add if they are in the same side.
- 6- We calculate the final value of (ΔH) from its value in the sub-equations.

Calculate (ΔH) for this reaction:

S is reactant and 2 mole

SO₃ is product and 2 mole

➤ The first step: We compare the sub-equations with the total equation and check the reactants, products and number of moles.

S is reactant and 1 mole so multiply by 2

2x



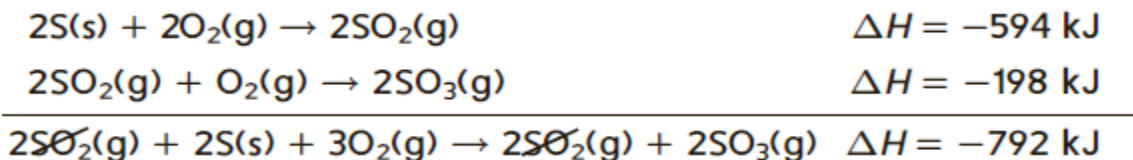
SO₃ reactants and 2 moles and in the overall equation in products and 2 moles so we reverse the equation



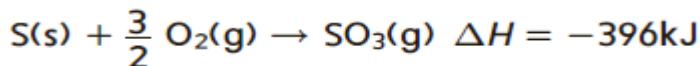
We reverse the equation and change the sign of ΔH .



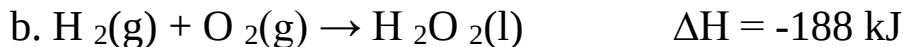
Step two: add the two equations 1, 2



- Thermochemical equations are usually written and balanced for one mole of product. Often, that means that **fractional coefficients must be used**.
- For example, the thermochemical equation for the reaction between sulfur and oxygen to form one mole of sulfur trioxide is the following.



1- Use thermochemical Equations a and b below to determine ΔH for the decomposition of hydrogen peroxide (H₂O₂).



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2- Use Equations a and b below to determine ΔH for the following reaction:



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3- Challenge ΔH for the following reaction is -1789 kJ. Use this and Equation a to determine ΔH for Equation b.



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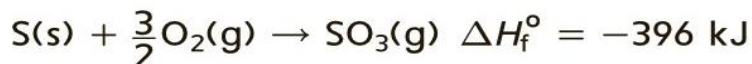
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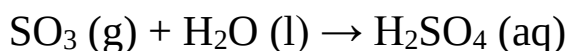
Standard Enthalpy (Heat) of Formation: (ΔH_f°)

Is the change in enthalpy that accompanies the formation of one mole of the compound in its standard state from its elements in their standard states.

- typical standard heat of formation reaction is the formation of one mole of SO_3 from its elements.



- SO_3 gas is a suffocating gas that produces acid rain when it mixes with atmospheric moisture and forms sulfuric acid, which increases the acidity of rain.

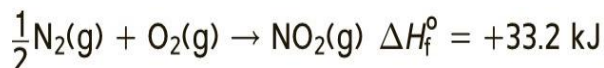


- **Acid rain damage:** destroys forests, destroys fisheries, and damages sculptures.

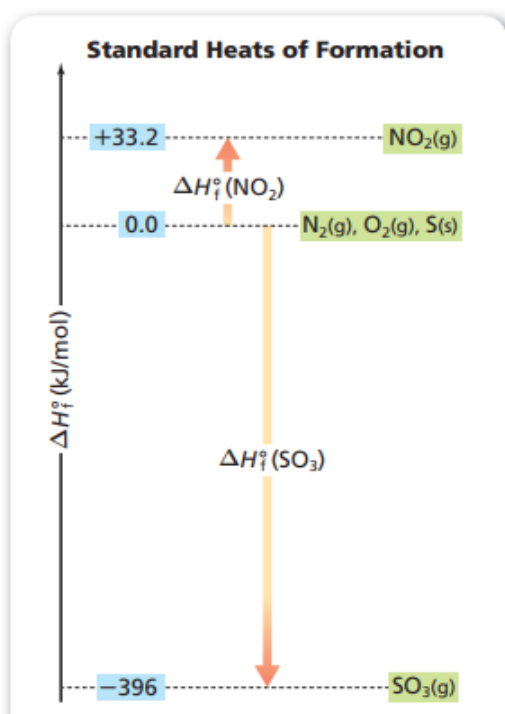
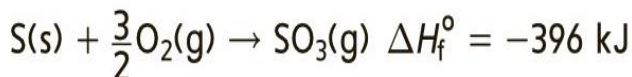


Where do standard heats of formation come from?

- The standard heat of formation (ΔH_f° for the elements = 0.0 KJ.)
- The standard heat of formation of the compounds has been experimentally calculated and organized into a higher and lower scale than the elements in their standard states.
- The formation temperature is higher than zero, which is positive, such as the nitrogen dioxide formation equation.



- The heat of formation value is +33.2 KJ, which means that the reaction is endothermic and the energy of the products is greater than the energy of the reactants (0.0 KJ) by +33.2 KJ.
- The heat of formation is less than zero, i.e. negative, such as the equation for the formation of sulfur trioxide.



- The value of the heat of formation is -396 KJ, which means that the reaction is exothermic, so the energy of the products is less than the energy of the reactants (0.0 KJ) by (-396 KJ).



■ **Figure 15.16** Sulfur hexafluoride is used to etch minute and sometimes intricate patterns on silicon wafers in the production semiconductor devices. Semiconductors are important components of modern electronic equipment, including computers, cell phones, and MP3 players.

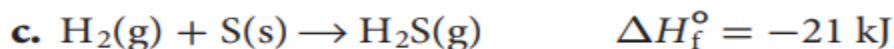
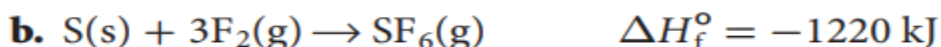
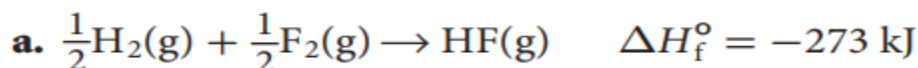
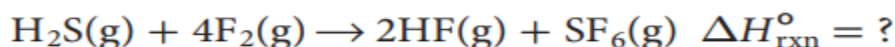
Table 15.5		
Standard Enthalpies of Formation		
Compound	Formation Equation	ΔH_f° (kJ/mol)
H ₂ S(g)	H ₂ (g) + S(s) → H ₂ S(g)	-21
HF(g)	$\frac{1}{2}$ H ₂ (g) + $\frac{1}{2}$ F ₂ (g) → HF(g)	-273
SO ₃ (g)	S(s) + $\frac{3}{2}$ O ₂ (g) → SO ₃ (g)	-396
SF ₆ (g)	S(s) + 3F ₂ (g) → SF ₆ (g)	-1220

The heat of formation and stability of compound :

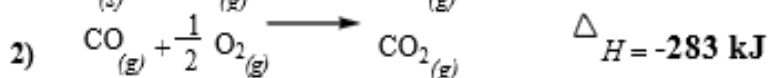
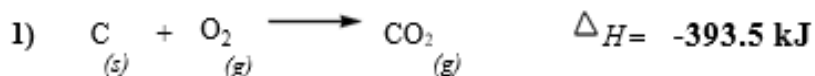
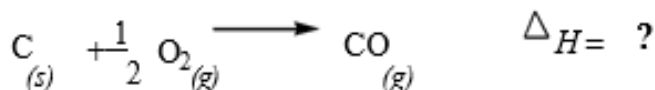
- When the heat of formation is more negative the stability increases, When the heat of formation is more positive the stability decreases.

Using standard enthalpies of formation :

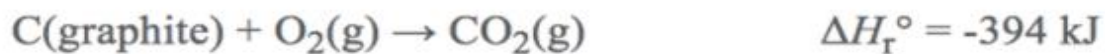
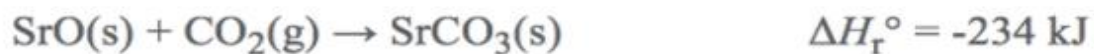
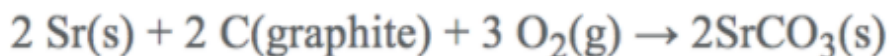
- Standard enthalpies of formation can be used to calculate the enthalpies of many reactions under standard conditions $\Delta H_{\text{rxn}}^\circ$ using Hess's law.



Calculate the enthalpy heat of formation of CO :



Calculate the enthalpy heat of formation of SrCO₃ :



The summation equation :

- The stepwise procedure you have just read about shows how standard heats of formation equations combine to produce the desired equation and its $\Delta H_{\text{rxn}}^\circ$.
- The procedure can be summed up in the following formula.

➤ Summation Equation

$$\Delta H_{\text{rxn}}^\circ = \Sigma \Delta H_f^\circ (\text{products}) - \Sigma \Delta H_f^\circ (\text{reactants})$$

$\Delta H_{\text{rxn}}^\circ$: represents the standard enthalpy of the reaction.

Σ represents the sum of the terms.

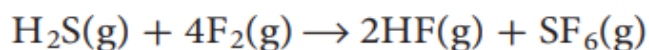
$\Delta H_f^\circ(\text{products})$: represent the standard enthalpies of formation of all the products

$\Delta H_f^\circ(\text{reactants})$: represent the standard enthalpies of formation of all the reactants.

$\Delta H_{\text{rxn}}^\circ$ is obtained by subtracting the sum of heats of formation of the reactants from the sum of the heats of formation of the products.

- You can see how this formula applies to the reaction between hydrogen sulfide and fluorine.

- $\text{HF(g)} \Delta H_f^\circ = -273 \text{ kJ}$, $\text{SF}_6(\text{g}) \Delta H_f^\circ = -1220 \text{ kJ}$, $\text{H}_2\text{S(g)} \Delta H_f^\circ = -21 \text{ kJ}$



$$\Delta H_{\text{rxn}}^\circ = [(2)\Delta H_f^\circ(\text{HF}) + \Delta H_f^\circ(\text{SF}_6)] - [\Delta H_f^\circ(\text{H}_2\text{S}) + (4)\Delta H_f^\circ(\text{F}_6)]$$

$$\Delta H_{\text{rxn}}^\circ = [(2)(-273 \text{ kJ}) + (-1220 \text{ kJ})] - [-21 \text{ kJ} + (4)(0.0 \text{ kJ})]$$

$$\Delta H_{\text{rxn}}^\circ = -1745 \text{ kJ}$$

- Use standard enthalpies of formation to calculate $\Delta H_{\text{rxn}}^\circ$ for the combustion of methane. $\Delta H_{\text{rxn}}^\circ = ?$



- $\Delta H_f^\circ(\text{CO}_2) = -394 \text{ kJ}$, $\Delta H_f^\circ(\text{H}_2\text{O}) = -286 \text{ kJ}$, $\Delta H_f^\circ(\text{CH}_4) = -75 \text{ kJ}$
 $\Delta H_f^\circ(\text{O}_2) = 0.0 \text{ kJ}$

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- Show how the sum of enthalpy of formation equations produces each of the following reactions. You do not need to look up and include H values.

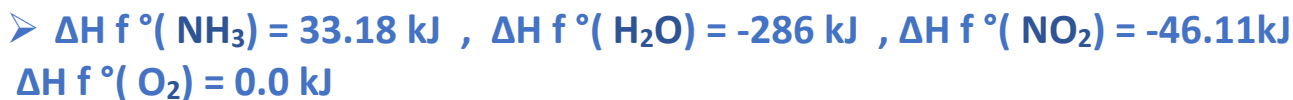


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- Use standard enthalpies of formation to calculate $\Delta H_{\text{rxn}}^\circ$ for the following reaction.



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- Determine $\Delta H_{\text{comb}}^\circ$ for butanoic acid,
 $\text{C}_3\text{H}_7\text{COOH}(\text{l}) + 5\text{O}_2(\text{g}) \rightarrow 4\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{l})$. Use data that given to you and the following equation. $4\text{C}(\text{s}) + 4\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{C}_3\text{H}_7\text{COOH}(\text{l})$ $\Delta H = -534 \text{ kJ}$



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➤ **Challenge** Two enthalpy of formation equations, a and b, combine to form the equation for the reaction of nitrogen oxide and oxygen.

The product of the reaction is nitrogen dioxide:



What is ΔH_f° for Equation b?

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➤ **Calculate** Use Hess's law to determine ΔH for the reaction



given the following reactions. Show your work.



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- Use standard enthalpies of formation to calculate $\Delta H_{\text{rxn}}^\circ$ for the following reaction.



- $\Delta H_f^\circ(\text{P}_4\text{O}_{10}) = -2984 \text{ kJ}$, $\Delta H_f^\circ(\text{P}_4\text{O}_6) = -1640.1 \text{ kJ}$, $\Delta H_f^\circ(\text{O}_2) = 0.0 \text{ kJ}$.

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- Use Hess's law and the following thermochemical equations to produce the thermochemical equation for the reaction $\text{C}(\text{s, diamond}) \rightarrow \text{C}(\text{s, graphite})$. What is ΔH for the reaction?

- a. $\text{C}(\text{s, graphite}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$ $\Delta H = -394 \text{ kJ}$

- b. $\text{C}(\text{s, diamond}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$ $\Delta H = -396 \text{ kJ}$

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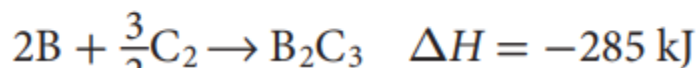
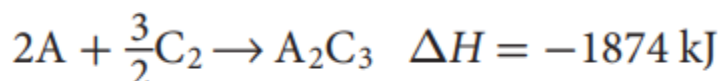
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- Use Hess's law and the changes in enthalpy for the following two generic reactions to calculate ΔH for the reaction: $2\text{A} + \text{B}_2\text{C}_3 \rightarrow 2\text{B} + \text{A}_2\text{C}_3$.



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➤ Determine if the reaction is a heat of combustion reaction, a heat of formation reaction, or not:

1-	$\text{C}_{(s)} + \text{O}_{2(g)} \longrightarrow \text{CO}_{2(g)} \quad \Delta H = -393.5 \text{ kJ}$	
2-	$2\text{H}_{2(g)} + \text{O}_{2(g)} \longrightarrow 2\text{H}_2\text{O}_{(l)} \quad \Delta H = -571.6 \text{ kJ}$	
3-	$\text{S}_{(s)} + \text{O}_{2(g)} \longrightarrow \text{SO}_{2(g)} \quad \Delta H = -296.8 \text{ kJ}$	
4-	$\text{N}_{2(g)} + 2\text{O}_{2(g)} \longrightarrow 2\text{NO}_{2(g)} \quad \Delta H = +66.4 \text{ kJ}$	
5-	$\text{CH}_{4(g)} + 2\text{O}_{2(g)} \longrightarrow \text{CO}_{2(g)} + 2\text{H}_2\text{O}_{(l)} \quad \Delta H = -891 \text{ kJ}$	
6-	$\text{C}_{(s)} + 2\text{H}_{2(g)} \longrightarrow \text{CH}_{4(g)} \quad \Delta H = -74.7 \text{ kJ}$	
7-	$\text{H}_{2(g)} + \frac{1}{2}\text{O}_{2(g)} \longrightarrow \text{H}_2\text{O}_{(l)} \quad \Delta H = -285.8 \text{ kJ}$	

➤ The heat of formation of the compound X is (-110 KJ) , the heat of formation of the only product of combustion of this compound is (-393.5KJ). what is the heat of combustion of the compound X ?

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➤ The heat of formation of the compound X is (- 612 KJ) , the heat of formation of the only product of combustion of this compound is (-671KJ). what is the heat of combustion of the compound X ?

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