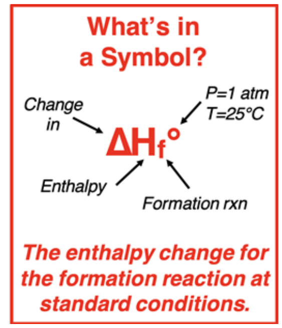
Thermochemistry **Name**____________________________________

Heat of Formation and Bond Enthalpy

Read from **Lesson 2 in** the **Chemistry Tutorial Section, Chapter 12** of **The Physics Classroom**:
Part c: Heat of Formation Part d: Bond Enthalpy and ΔH

Part 1: Heat of Formation

In chemistry, the **heat of formation** (also called **enthalpy of formation**, symbolized as $\Delta H_f°$) is the amount of energy released or absorbed when **one mole of a compound** is formed from its **elements in their standard states** under standard conditions (25°C and 1 atmosphere of pressure).



Think of it like this: if you could build a compound from scratch using the pure elements from the periodic table, the heat of formation tells you how much energy that process would take or give off.

For example:
The formation of water from hydrogen and oxygen:

$$H_2\,(g) + \tfrac{1}{2}\,O_2\,(g) \rightarrow H_2O\,(l) \qquad \Delta H_f° = -285.8\ \text{kJ/mol}$$

The formation of butane from carbon and hydrogen:

$$4\,C\,(s) + 5H_2\,(g) \rightarrow C_4H_{10}(g) \qquad \Delta H_f° = -124.7\ \text{kJ/mol}$$

Other compounds' heat of formation values can be found in the reference section of the Chemistry Tutorial. This concept is fundamental to understanding chemical energy changes and plays a crucial role in calculating reaction enthalpies using **Hess's Law**.

> The enthalpy change of any reaction is the sum of the $\Delta H_f°$ values of all the product compounds minus the sum of the $\Delta H_f°$ of all reactant compounds. In equation form:
>
> $$\Delta H_{reaction} = \sum \Delta H_f°_{products} - \sum \Delta H_f°_{reactants}$$

Example:

Write a reaction for the combustion of butane, C_4H_{10}.: $2\,C_4H_{10}\,(g) + 13\,O_2\,(g) \rightarrow 8\,CO_2\,(g) + 10\,H_2O\,(g)$

Calculate the $\Delta H_{combustion}$ for the combustion reaction. $\Delta H_{reaction} = \sum \Delta H_f°_{products} - \sum \Delta H_f°_{reactants}$
(look up heats of formation from the reference table)
$\Delta H_{combustion} = (8 * H_f°\text{of } CO_2 + 10 * H_f° \text{ of } H_2O) - (2 * H_f° \text{ of } C_4H_{10} + 13 * H_f° \text{ of } O_2)$
$\Delta H_{combustion} = [(8 * -393.5\ \text{kJ}) + (10 * -241.8\ \text{kJ})] - [(2 * -124.7\ \text{kJ/mol} + (13 * 0\ \text{kJ})] = -5566\ \text{kJ} - (-249.4) = -5317\ \text{kJ}$ (for 2 moles)
$\Delta H_{combustion}$ for butane = 2658 kJ/mol

Questions

1. Identify the standard state (solid, liquid or gas) for the following elements:

a. carbon	d. argon
b. nitrogen	e. mercury
c. bromine	f. iron

2. Use the heat of formation values found in the reference section of the Chemistry Tutorial to find the standard enthalpy of formation, $\Delta H_f°$ for the following:

a. $NaCl(s)$	e. $Fe_2O_3(s)$
b. $NH_3(g)$	f. $Bi(s)$
c. $H_3PO_4(l)$	g. $UF_6(s)$
d. $O_3(g)$	h. $CCl_2F_2(g)$

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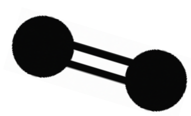
3. Use the heat of formation values found in the reference section of the Chemistry Tutorial to determine the enthalpy change for the following reactions.
   a. $4\ NH_3\ (g) + 7\ O_2(g) \rightarrow 4\ NO_2\ (g) + 6\ H_2O(g)$

   b. $HNO_3(aq) + NaOH(s) \rightarrow NaNO_3(aq) + H_2O(l)$

4. The combustion reaction of acetone is $C_3H_6O\ (l) + 4\ O_2\ (g) \rightarrow 3\ CO_2(g) + 3\ H_2O\ (l)$ $\Delta H_{combustion}$ is −1790 kJ/mol. Use this and the heat of formation values found in the reference section of the Chemistry Tutorial to determine the heat of formation of acetone.

Part 2: Bond Enthalpy

Bond enthalpy, or **bond dissociation energy**, is the energy required to break **one mole** of a specific bond in a gas-phase molecule. It reflects bond strength—**stronger bonds have higher bond enthalpies**, as they require more energy to break. Chemists use bond enthalpies to predict energy changes in chemical reactions. By comparing bond energies in reactants and products, we can estimate whether a reaction absorbs or releases energy.



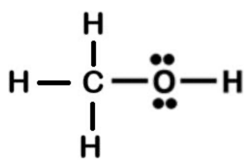
Using Bond Enthalpy to Calculate ΔH Values

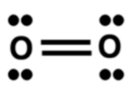
1. **Conduct a Bond Inventory**
   Identify the bonds being **broken** and **formed** in the reaction. Determine the quantity of each bond type. Use Lewis electron dot diagrams to classify bonds as single, double, or triple.
2. **Look Up Bond Enthalpy Values**
   Refer to the Bond Enthalpy Table found in the reference section of the Chemistry Tutorial to find the enthalpy values for each bond being broken and formed.
3. **Calculate the ΔH for Bond Breaking**
   For each bond broken, multiply the number of bonds by the bond enthalpy. Sum all values to determine ΔH for the bond-breaking steps. This value will be positive, as breaking bonds requires energy.

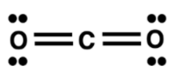
4. **Calculate the ΔH for Bond Forming**
   Bond formation is an exothermic process. The ΔH for forming bonds is negative—simply write the bond enthalpy with a negative sign. Multiply the number of bonds formed by their respective bond enthalpy values, then sum these negative values to determine ΔH for bond formation.
5. **Determine the Overall Enthalpy Change**
   Add the ΔH for bond breaking (positive) to the ΔH for bond forming (negative). The result is the overall enthalpy change (ΔH), indicating whether the reaction is endothermic or exothermic.

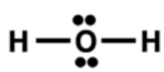
For example: Refer to the Bond Enthalpy Table found in the reference section of the Chemistry Tutorial to find the enthalpy values for each bond being broken and formed for this reaction:

$$2\ CH_3OH\ (l)\ +3\ O_{2.}\ (g) \rightarrow 2\ CO_2\ (g)\ +\ 4\ H_2O\ (g)$$

1. **Conduct a Bond Inventory:** Identify the bonds being **broken** and **formed** in the reaction
   Bonds broken:
   CH_3OH  has 3 C-H bonds, 1 C-O bond, and 1 O-H bond

   O_2  has 1 O=O bond

   Bonds formed:
   CO_2  has 2 C=O bonds

   H_2O  has 2 O-H bonds

2 **Look Up Bond Enthalpy Values** on the Bond Enthalpy Table
   C-H (413 kJ) C-O (358 kJ) O-H (467 kJ) O=O (495 kJ) C=O (745 kJ)

3. **Calculate the ΔH for Bond Breaking**
   CH_3OH: 3 C-H (3*413 kJ) + C-O (358 kJ) +O-H (467 kJ) = 2064 kJ — 2 CH_3OH (4128 kJ)
   O_2 : O=O (495 kJ) — 3 O_2 (1485 kJ)

4. **Calculate the ΔH for Bond Forming**
   CO_2 : 2 C=O (2*745 kJ) = 1490 kJ — 2 CO_2 (2980 kJ)
   H_2O : 2 O-H (2*467 kJ) = 934 kJ — 4 H_2O (3736 kJ)

5. **Determine the Overall Enthalpy Change**
   ΔH_{rxn} = (+bonds breaking) + (–bonds forming) = (4128 kJ + 1485 kJ) + (-2980 kJ – 3736 kJ) = -1103 kJ
   ↑ reactants ↑ ↑ products ↑

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Questions

Use the Bond Enthalpy Table found in the reference section of the Chemistry Tutorial to determine the heat of reaction for the following reactions.

1. $N_2H_4 \rightarrow N_2 + 2\ H_2$

2. $2\ H_2O_2 \rightarrow 2\ H_2O + O_2$

3. $CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$

4. $C_3H_6O + 4\ O_2 \rightarrow 3\ CO_2 + 3\ H_2O$