CHEM1101 Worksheet 10: Enthalpy of Reaction ($\Delta_{rxn}H$)

Model 1: Endothermic and Exothermic Processes

Breaking bonds requires energy to pull the atoms apart: bond breaking is endothermic ($\Delta H > 0$). When bonds are formed, energy is released – precisely the same amount of energy which would be required to break those bonds: bond making is exothermic ($\Delta H < 0$).

In most chemical reactions, bonds are broken and made. Whether a reaction is endothermic or exothermic depends on the energy required to perform the *changes* in the bonding.

Critical thinking questions

1. Are the following reactions exothermic or endothermic?

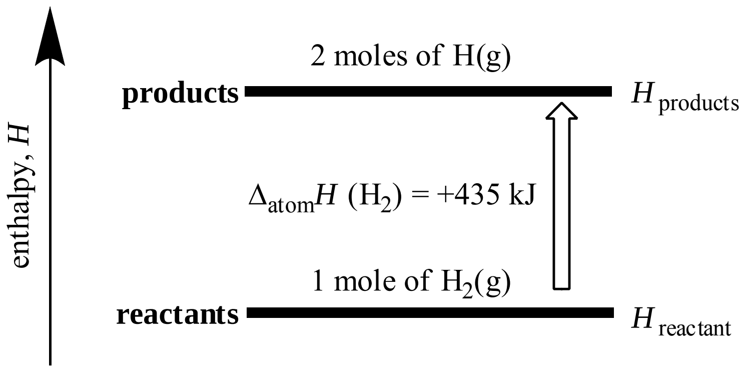
   (a) $H_2(g) \rightarrow 2H(g)$ (b) $Cl_2(g) \rightarrow 2Cl(g)$

   (c) $H(g) + Cl(g) \rightarrow HCl(g)$

Model 2: Enthalpy of Atomization ($\Delta_{atom}H$) and Enthalpy of Atom Combination ($\Delta_{ac}H$)

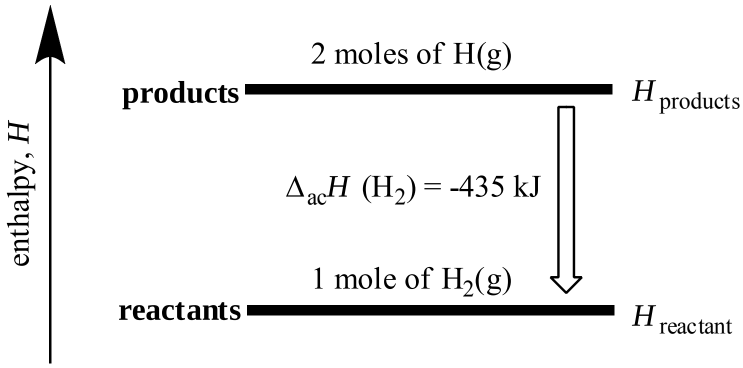
When a mole of a compound is broken apart into its constituent gas phase atoms, energy is consumed and the energy change is called the enthalpy of atomization ($\Delta_{atom}H$):

$$\Delta_{atom}H = H(\text{products}) - H(\text{reactants})$$
$$= H(\text{atoms}) - H(\text{compound}) \quad (1)$$



When a mole of a compound is made from its constituent gas phase atoms, energy is released and the energy change is called the enthalpy of atom combination ($\Delta_{ac}H$):

$$\Delta_{ac}H = H(\text{products}) - H(\text{reactants})$$
$$= H(\text{compound}) - H(\text{atoms}) \quad (2)$$



Critical thinking questions

1. What is the relationship between $\Delta_{atom}H$ and $\Delta_{ac}H$ for a *compound* like H_2?

2. What is the value of ΔH for the overall process of separating one mole of $H_2(g)$ into its constituent atoms and then reforming one mole of $H_2(g)$?

Model 3: Enthalpy of Reaction using $\Delta_{atom}H$ and $\Delta_{ac}H$

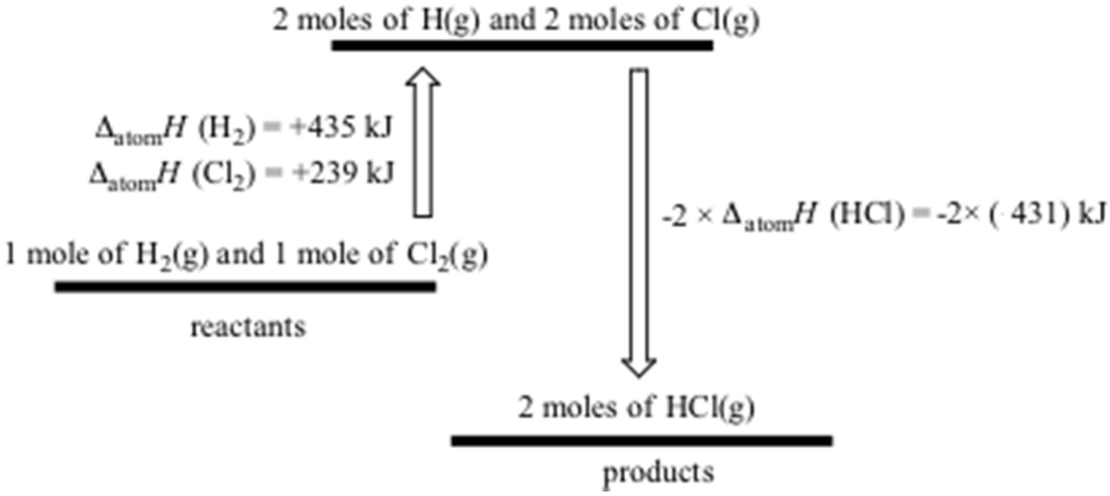
To determine the overall value of ΔH for a reaction, we can imagine the reaction taking place by:

(i) breaking apart all of the reactant molecules into their constituent atoms: $\Delta_{atom}H$ (reactants)

(ii) reassembling or combining these atoms into the product molecules: $\Delta_{ac}H$ (products)

The overall enthalpy of the reaction is then the sum of these parts:

$$\Delta_{rxn}H = \Delta_{atom}H \text{ (reactants)} + \Delta_{ac}H \text{ (products)} \qquad (3)$$



Critical thinking questions

1. Why is the ΔH associated with the upward arrow in Model 3 a *positive* number?

2. Why is the ΔH associated with the downward arrow in Model 4 a *negative* number?

3. What is the value of ΔH for the overall reaction in Model 3?

4. Using your answer to Q2, rewrite the equation below so that it involves only $\Delta_{ac}H$ (reactants) and $\Delta_{ac}H$ (products).

   $\Delta_{rxn}H = \Delta_{atom}H$ (reactants) + $\Delta_{ac}H$ (products) =

5. Using your answer to Q2, rewrite the equation below so that it involves only $\Delta_{atom}H$ (reactants) and $\Delta_{atom}H$ (products).

   $\Delta_{rxn}H = \Delta_{atom}H$ (reactants) + $\Delta_{ac}H$ (products) =

6. If $\Delta_{ac}H$ for the reactants is more negative than $\Delta_{ac}H$ for the products in a chemical reaction, will $\Delta_{rxn}H$ be positive or negative? Explain your reasoning.

Model 4: Enthalpy of Reaction using $\Delta_f H$

In Model 2, you developed a way of working out the value of enthalpy change for a reaction from the values of enthalpy of *atom* combination for the reactants and products. From Q7:

$$\Delta_{rxn}H = \Delta_{ac}H \text{ (products)} - \Delta_{ac}H \text{ (reactants)} \qquad (4)$$

An alternative is to use the enthalpy change of formation of a compound ($\Delta_f H$) from its *elements* in their naturally occurring forms. At room temperature and pressure, these forms are called the **standard states** of the elements and include, for example, graphite for carbon and $O_2(g)$ for oxygen.

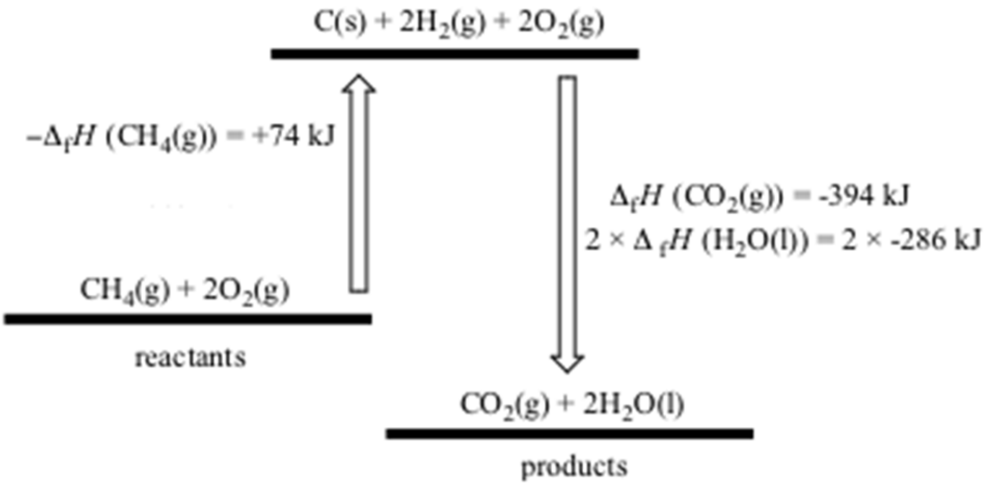
Using this method, the equation for the enthalpy of reaction becomes:

$$\Delta_{rxn}H° = \Delta_f H° \text{ (products)} – \Delta_f H° \text{ (reactants)} \qquad (5)$$

The enthalpy of formation of $CO_2(g)$ is then the energy change for its formation from graphite and $O_2(g)$:

$$C(s) + ½\ O_2(g) \rightarrow CO_2(g)$$

The enthalpy change for the combustion of methane is represented on the energy level diagram below. On the left, $CH_4(g)$ and $O_2(g)$ are broken up into their elements in the standard states, graphite ($C(s)$), $H_2(g)$ and $O_2(g)$. This is the *reverse* of their formation so the energy required is $–\Delta_f H°$ (reactants). On the right, $CO_2(g)$ and $H_2O(g)$ are formed from the same elements in the same states so the energy change is $+ \Delta_f H°$ (products).



Critical thinking questions

1. Why are $\Delta_f H°$ ($O_2(g)$) and $\Delta_f H°$ ($H_2(g)$) both equal to 0 kJ? (*Hint*: what is the reaction in each case?)

2. What is $\Delta_{rxn}H°$ for the reaction $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$?

3. Use equation (5) and the data below to calculate $\Delta_{rxn}H°$ for the reaction $MgO(s) + CO_2(g) \rightarrow MgCO_3(s)$.
$\Delta_f H°$: $MgO(s) = -602$ kJ mol^{-1}, $CO_2(g) = -394$ kJ mol^{-1} and $MgCO_3(s) = -1096$ kJ mol^{-1}

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Marks 3

- Pentane, $CH_3CH_2CH_2CH_2CH_3$, burns completely in oxygen to form CO_2(g) and H_2O(g). Use the atomization enthalpies given below to calculate the enthalpy change for this process.

	$\Delta_{atom}H$ (kJ mol^{-1})		$\Delta_{atom}H$ (kJ mol^{-1})
pentane	6352	CO_2	1608
O_2	498	H_2O	926

Answer:

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Marks 5

- The current "petrochemical economy" is based on the combustion of fossil fuels, of which octane is a typical example.

$$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(l)$$

Calculate the heat of combustion of octane using the supplied heat of formation data.

Data: C_8H_{18}(l): –249.9 kJ mol^{-1}; CO_2(g): –393.5 kJ mol^{-1}; H_2O(l): –285.8 kJ mol^{-1}

Answer: