

UWO **CHEM 1302**

Winter 2025, Chapter 4 Notes



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4.1 Introduction to Enthalpy Calculations

4.1.1

Enthalpy of Reaction (ΔH°_{rxn})

Enthalpy (H): is a measure of the energy associated with breaking or forming bonds

- Breaking bonds (requires/releases) requires energy
 - Therefore, this is (endothermic/exothermic) endothermic
 - *Think: you need to be strong to break bonds!
- Forming bonds (requires/releases) releases energy
 - Therefore, this is (endothermic/exothermic) exothermic
 - *Think: bonds want to form if they are stable and lower in energy!

WIZE CONCEPT

In summary: breaking bonds requires energy & forming bonds releases energy!

We will soon look at the different ways to calculate ΔH°_{rxn} :

The Heats of Formation Method (ΔH°_f)

Average Bond Enthalpy Method (BDE)

Hess' Law of Formation Method

Example: Calculating Enthalpy of a Reaction



1. How much heat is produced when 72g of water gas is produced?

$$m = 72 \text{ g} \quad M = 18 \text{ g/mol} \quad 2 \text{ mols H}_2\text{O} \text{ heat release } 572 \text{ kJ,}$$

$$n = \frac{72}{18} = 4 \text{ mol} \quad 4 \text{ mols H}_2\text{O will release heat of } 1144 \text{ kJ.}$$

2. Is the reaction endothermic or exothermic?

$\therefore \Delta H$ is negative \rightarrow heat is released

\therefore reaction is exothermic



3. We are told that $\text{O}_2(\text{g})$ is the limiting reagent of this reaction. How many moles of $\text{O}_2(\text{g})$ are used if the reaction releases 286 kJ of heat?

$$n_{\text{O}_2} = \frac{\Delta H_{\text{ad}}}{\Delta H} = \frac{-286 \text{ kJ}}{-572 \text{ kJ/mol}} = 0.5 \text{ mol}$$

- ∴ The reaction has 1 mol O_2 leading to 572 kJ of heat being released.
- ∴ $\frac{1}{2}$ mol of O_2 would lead to 286 kJ of heat being released

4.2

Hess' Law Method

4.2.1

Hess' Law

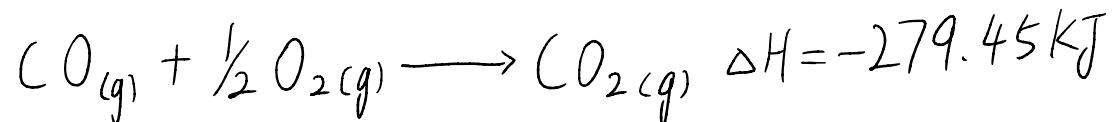
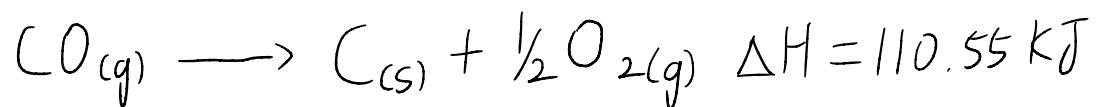
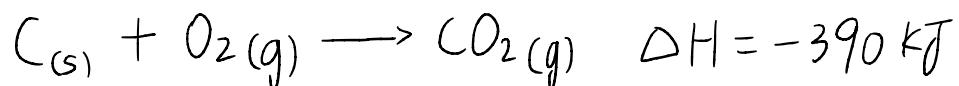
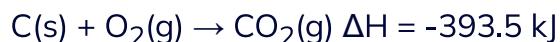
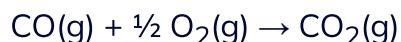
If a reaction is carried out in a series of steps, ΔH for the overall reaction can be found from the sum of the enthalpy changes of the individual steps.

This is because enthalpy (H) is a state function and when it changes, it does not depend on the pathway taken!

$$\Delta H_{rxn} = \Delta H_1 + \Delta H_2 + \Delta H_3 + \dots$$

Example:

What is the ΔH for the reaction?





WIZE CONCEPT

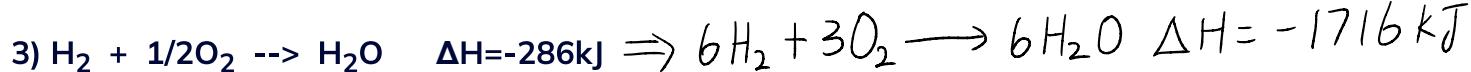
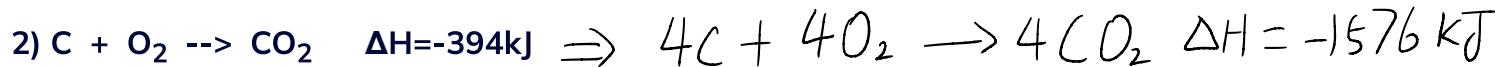
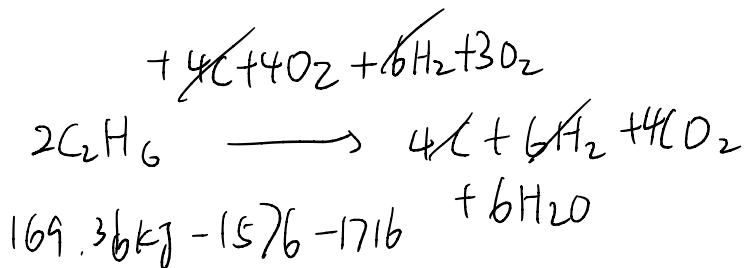
The two things you need to remember for this method are:

- 1) Whenever we want to look at a **reaction in reverse**, we must **multiply ΔH by -1**
- 2) Whenever we want to **multiply a reaction by a coefficient**, we must **multiply the ΔH of that step by that same coefficient!**

4.2.2

Practice: Calculating the Enthalpy of a Reaction Using Hess' Law

What is the ΔH_{rxn} for the following reaction?



-3200 kJ

-3000 kJ

-1800 kJ

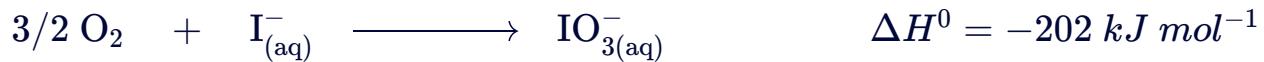
1400 kJ

$$\Delta H_{rxn}^\circ = (169.36) + (-1576) + (-1716)$$

$$\Delta H_{rxn}^\circ = -3122\text{ kJ}$$

4.2.3

Given that the $\Delta_f H^0$ of ozone $O_{3(g)}$ is $142.67 \text{ kJ mol}^{-1}$ and the reaction shown below, calculate the enthalpy for the unknown reaction

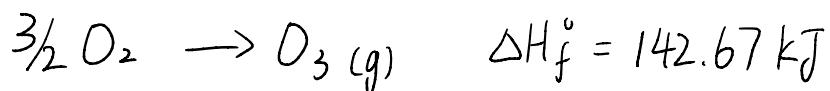
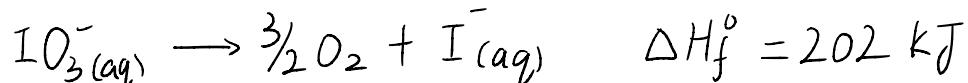


106 kJ/mol

196 kJ/mol

344 kJ/mol

426 kJ/mol



$$\Delta H_f^\circ = 202 + 142.67 = 344.67 \text{ kJ}$$

$\therefore \Delta H_m^\circ$ is endothermic (absorb heat)

4.3

Average Bond Enthalpy Method

4.3.1

Bond Energies

Enthalpy (H) is the energy stored in bonds.

Bond dissociation energy (BDE) is the energy needed to break a chemical bond.

$$\Delta H_{rxn} = [(\sum nH_{bonds\ broken}) - (\sum nH_{bonds\ formed})]$$

Or (an easier way to think about it):

$$\boxed{\Delta H_{rxn} = [(\sum nBDE_{reactants}) - (\sum nBDE_{products})]}$$

ΔH_{rxn} = enthalpy change for the reaction (kJ/mol)

BDE = bond energy per mole of bonds (kJ/mol), always positive

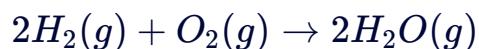
$$\Delta H_{rxn} = [(\sum nBDE_{reactants}) - (\sum nBDE_{products})]$$

Bond Energies (kJ/mol)					
Bond	Bond Energy	Bond	Bond Energy	Bond	Bond Energy
H-H	436	C-S	260	F-Cl	255
H-C	415	C-Cl	330	F-Br	235
H-N	390	C-Br	275	Si-Si	230
H-O	464	C-I	240	Si-P	215
H-F	569	N-N	160	Si-S	225
H-Si	395	N=N	418	Si-Cl	359
H-P	320	N≡N	946	Si-Br	290
H-S	340	N-O	200	Si-I	215
H-Cl	432	N-F	270	P-P	215
H-Br	370	N-P	210	P-S	230
H-I	295	N-Cl	200	P-Cl	330
C-C	345	N-Br	245	P-Br	270
C=C	611	O-O	140	P-I	215
C≡C	837	O=O	498	S-S	215

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Example: Using Bond Dissociation Energies to Calculate Enthalpy of a Reaction

What is ΔH_{rxn} for the following reaction?



Bond	Bond Energy (kJ/mol)
H-H	436
O=O	499
O-H	463

$$\begin{aligned}\Delta H_{rxn}^{\circ} &= [(\sum n BDE_{\text{reactant}}) - (\sum n BDE_{\text{products}})] \\ &= (2 \times 439 + 499) - 4 \times 463 \\ &= -481 \text{ kJ}\end{aligned}$$

Practice: Calculating the Enthalpy of a Reaction Using Bond Energies

Use the table of average bond energies to calculate the ΔH for the following reaction:



<u>Bond</u>	<u>Bond Energy (kJ/mol)</u>
H – H	432
H – C	413
H – I	295
I – I	149
C – I	240
C – C	347
C = C	614
C ≡ C	839

A) -601 kJ

B) +601 kJ

C) -323 kJ

D) +323 kJ

$$\begin{aligned}
 \Delta H &= \sum \text{BDE}_{(\text{Reactants})} - \sum \text{BDE}_{(\text{Products})} \\
 &= (893 + 2 \times 413) - (614 + 4 \times 413) \\
 &= 1665 - 2266 \\
 &= -601 \text{ kJ}
 \end{aligned}$$

$\therefore \Delta H$ is negative

\therefore the rxn is exothermic.