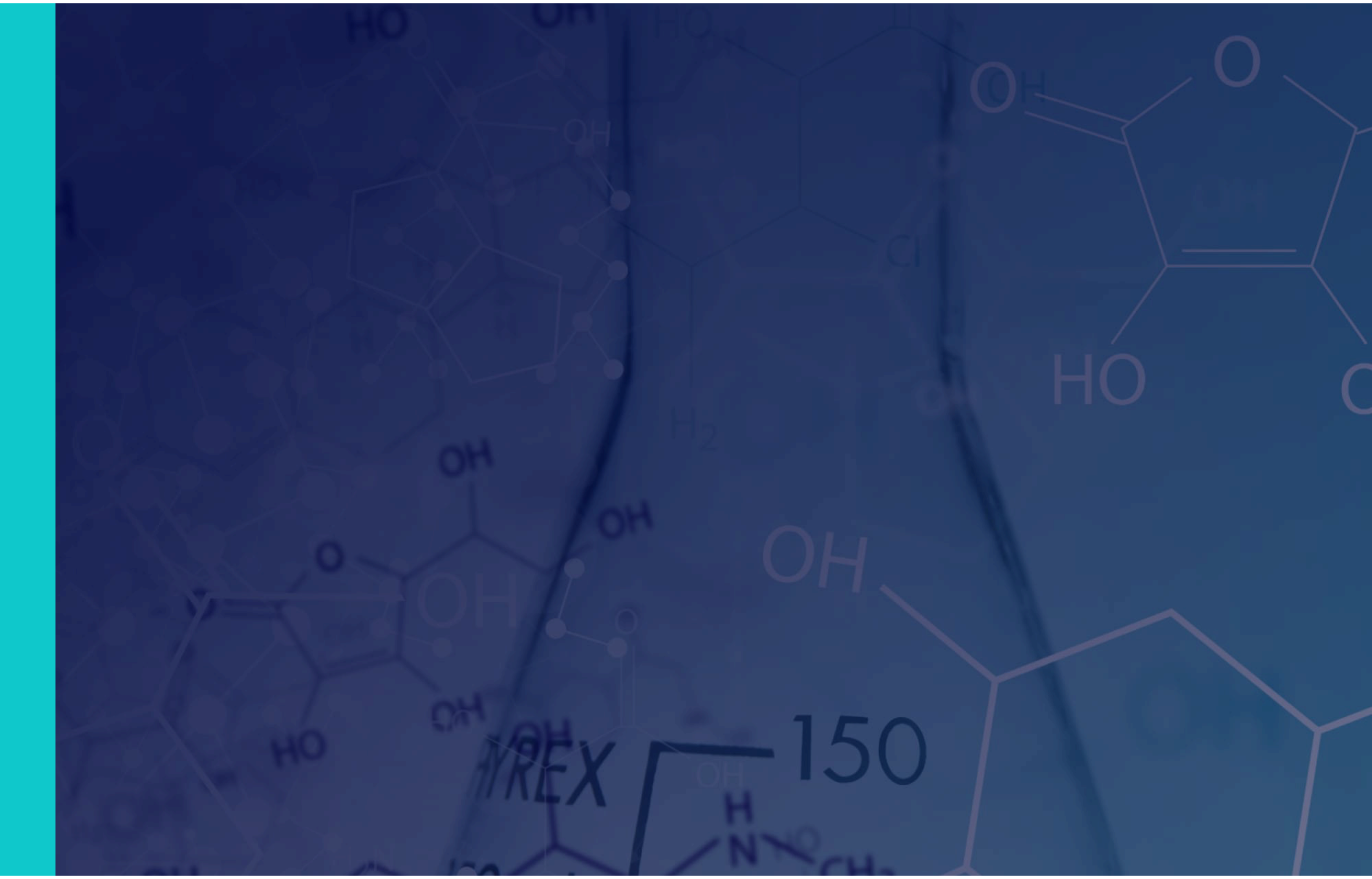

UWO

CHEM 1302

Winter 2025, Chapter 4 Notes

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4.1 Introduction to Enthalpy Calculations

4.1.1

Enthalpy of Reaction ($\Delta H^{o}rxn$)

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

- Breaking bonds (requires/releases) ________ energy
    - Therefore, this is (endothermic/exothermic) ____________
    - *Think: you need to be strong to break bonds!
- Forming bonds (requires/releases) ________ energy
    - Therefore, this is (endothermic/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 $\Delta H^{o}rxn$:

The Heats of Formation Method ($\Delta H^{o}f$)

Average Bond Enthalpy Method (BDE)

Hess' Law of Formation Method

Example: Calculating Enthalpy of a Reaction

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ $\Delta H = -572 kJ/mol$

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

2. Is the reaction endothermic or exothermic?

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l) \quad \Delta H = -572 kJ/mol$$

3. We are told that $O_2(g)$ is the limiting reagent of this reaction. How many moles of $O_2(g)$ are used if the reaction releases 286kJ of heat?

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!

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

Example:
What is the ΔH for the reaction?

$$CO(g) + \frac{1}{2} O_2(g) \rightarrow CO_2(g)$$

$$C(s) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -393.5 \text{ kJ}$$

$$2C(s) + O_2(g) \rightarrow 2CO(g) \quad \Delta H = -221.1 \text{ kJ}$$

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 ΔHrxn for the following reaction?

$2C_2H_6(g)$ + $7O_2(g)$ --> $4CO_2(g)$ + $6H_2O(g)$

1) 2C + $3H_2$ --> C_2H_6 ΔH=-84.68kJ

2) C + O_2 --> CO_2 ΔH=-394kJ

3) H_2 + $1/2O_2$ --> H_2O ΔH=-286kJ

- -3200 kJ
- -3000 kJ
- -1800 kJ
- 1400 kJ

4.2.3

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

$$3/2\ O_2 + I^-_{(aq)} \longrightarrow IO^-_{3(aq)} \qquad \Delta H^0 = -202\ kJ\ mol^{-1}$$

$$IO^-_{3(aq)} \longrightarrow I^-_{(aq)} + O_{3(g)} \qquad \Delta H^0 = ???$$

- 106 kJ/mol
- 196 kJ/mol
- 344 kJ/mol
- 426 kJ/mol

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.

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

Or (an easier way to think about it):

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

$\Delta H_{rxn} =$ enthalpy change for the reaction (kJ/mol)

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

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

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		$\mathbf{N=N}$	418		Si–Cl	359
H–P	320		$\mathbf{N\equiv N}$	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
$\mathbf{C=C}$	611		O–O	140		P–I	215
$\mathbf{C\equiv C}$	837		$\mathbf{O=O}$	498		S–S	215

4.3.2

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

What is ΔHrxn for the following reaction?

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

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

4.3.3

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:

$$H-C\equiv C-H \rightarrow H_2C=CH_2$$

Bond	Bond Energy (kJ/mol)
$H-H$	432
$H-C$	413
$H-I$	295
$I-I$	149
$C-I$	240
$C-C$	347
$C=C$	614
$C\equiv C$	839

A) -601 kJ

B) +601 kJ

C) -323 kJ

D) +323 kJ