Chapter 6

Reaction Kinetics

The rate law for a reaction describes how the reaction is affected by reactant concentration.

$$\text{rate} = k\,[A]^m\ [B]^n$$

- must determine values of exponents (m, n) and rate constant (k) experimentally

Consider the reaction:

$$2\,NO + 2\,H_2 \rightarrow N_2 + 2\,H_2O$$

a) Write the rate law equation for this reaction.

b) What is the overall order of this reaction?

c) Identify the effect on overall rate if the initial [NO] were tripled, and the initial $[H_2]$ were quadrupled.

d) Calculate a value for the rate constant, k, at this temperature.

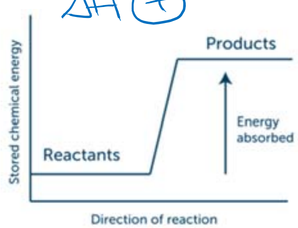
Run	Initial [NO] (mol/L)	Initial $[H_2]$ (mol/L)	Initial rate (mol/Ls)
1	0.400	0.100	1.10×10^{-5}
2	0.400	0.200	2.20×10^{-5}
3	0.800	0.200	8.80×10^{-5}

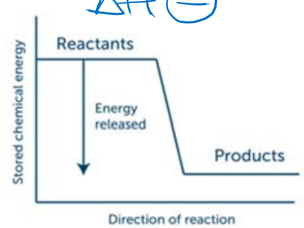
Chapter 5

Thermochemistry

Enthalpy (H) is the total energy content of a system.

- Changes (physical, chemical, nuclear) all result in a change in enthalpy, ΔH.
  - Endothermic vs Exothermic





- Can't measure ΔH directly.
- Can estimate ΔH:

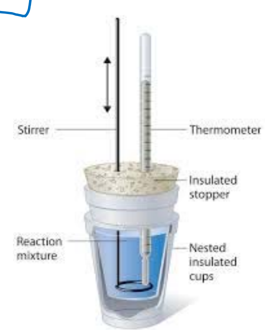
  1. **Experimentally** – Using experimental methods
     A. Calorimetry (5.2)

  2. **Theoretically** – Using calculations
     B. Average bond energies (5.3)
     C. Hess's law (5.4)
     D. Standard enthalpies of formation (5.5)

A. Calorimetry

Important equation:

$$q = m \ c \ \Delta T$$



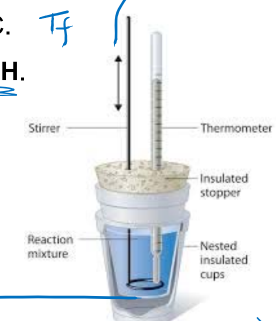
50.0 mL of 1.0 mol/L NaOH is neutralized by 50.0 mL of 1.0 mol/L HCl.

The initial temperature of both solutions before mixing is 18.0°C.

The maximum temperature reached after mixing is 24.6°C.

Calculate the molar enthalpy of neutralization for NaOH.

Questions to ask yourself:
1. What is the system in this reaction? The surroundings?
2. Is this reaction endothermic or exothermic? How can you tell?
3. How is q_{system} related to ΔH_{system}?



B. Average Bond Energies

Important equation:

Reactants — Product

$$\Delta H = \text{[energy required to break bonds]} - \text{[energy released when forming new bonds]}$$

$$\Delta H = \Sigma n \times D \text{ (bonds broken)} - \Sigma n \times D \text{ (bonds formed)}$$

(bond energy (kJ/mol))

Using the bond energies in Table 1 (pg. 307), calculate the enthalpy change for the reaction in which carbon monoxide and steam combine to produce carbon dioxide and hydrogen gases:

$$CO + H_2O \rightarrow CO_2 + H_2$$

C≡O — H–O–H — O=C=O — H–H

1 mol C≡O — 2 mol O–H — 2 mol C=O — 1 mol H–H

1072 kJ/mol — 467 kJ/mol — 799 kJ/mol — 432 kJ/mol

Questions to ask yourself:

1. What are the Lewis structures for these molecules? What types of bonds (single, double, triple) are involved?
2. How many bonds are involved per reactant/product?

C. Hess' Law

Concept:

- For a reaction that can occur in elementary steps, the sum of the enthalpy changes of all the steps will equal the overall enthalpy change for the reaction.

Strategy:

- Add up the given equations to achieve the "target" equation
  - Cancel entities that appear on both sides of the equation
  - Combine entities that appear on the same side of the equation
- Whatever you do to the elementary equation, do the same to the ΔH value:
  - Reversing the equation → Reverse the sign of ΔH
  - Multiply equation by a coefficient → Multiply ΔH by the same coefficient

Consider the following equations:

(1) $4\,P + 3\,O_2 \rightarrow P_4O_6$ $\quad \Delta H = -1640$ kJ

(2) $P_4O_{10} \rightarrow P_4O_6 + 2\,O_2$ $\quad \Delta H = 1344$ kJ) × −1

Calculate ΔH for the reaction represented by the equation:

$$4\,P + 5\,O_2 \rightarrow P_4O_{10}$$

		ΔH
(1)	$4P + 3O_2 \rightarrow P_4O_6$	−1640 kJ
(2)×−1	$P_4O_6 + 2O_2 \rightarrow P_4O_{10}$	−1344 kJ
	$4P + 5O_2 \rightarrow P_4O_{10}$	−2984 kJ

D. Standard Enthalpies of Formation

Standard Enthalpy of Formation ($\Delta H°_f$) – The change in enthalpy that occurs during the formation of 1 mol of a compound from its elements, in their standard states

- Table 1 pg. 320 (compounds)
- Equals zero for elements in their standard states

Important equation:

$$\Delta H°_{reaction} = (\Sigma\, n_{products}\, \Delta H°_{products}) - (\Sigma\, n_{reactants}\, \Delta H°_{reactants})$$

Calculate the standard enthalpy of combustion for methanol, CH_3OH, using standard enthalpy of formation values (Pg. 320). Assume the only products are carbon dioxide, and liquid water. $H_2O(l)$

$$2\,CH_3OH + 3O_2 \rightarrow 2\,CO_2 + 4\,H_2O(l)$$

-239.1   0   -393.5   -285.8

$$\Delta H°_{rxn} = (\Sigma\, n_{prod} \times \Delta H°_{f\,prod}) - (\Sigma\, n_{react.} \times \Delta H°_{f.\,react})$$

Questions to ask yourself:

1. What is the balanced chemical equation for this reaction?
2. Are all entities in their standard states?