



CHE1031 Module 8 summary: Thermochemistry

8.1: Energy basics

Energy is the capacity to do work (applying a force to move matter). **Kinetic energy (KE)** is the energy of motion; **potential energy** is energy due to relative position, composition, or condition. When energy is converted from one form into another, energy is neither created nor destroyed (**law of conservation of energy or first law of thermodynamics**). Matter has **thermal energy** due to the KE of its molecules and **temperature** that corresponds to the average KE of its molecules. **Heat** is energy that is transferred between objects at different temperatures; it **flows** from a high to a low temperature. Chemical and physical processes can absorb heat (**endothermic**) or release heat (**exothermic**). The SI unit of energy, heat, and work is the **joule (J)**. **Specific heat** and **heat capacity** are measures of the energy needed to change the temperature of a substance or object. The amount of heat absorbed or released by a substance depends directly on the type of substance, its mass, and the temperature change it undergoes.

8.2: Calorimetry

Calorimetry is used to measure the amount of thermal energy transferred in a chemical or physical process. This requires careful measurement of the temperature change that occurs during the process and the masses of the **system and surroundings**. These measured quantities are then used to compute the amount of heat produced or consumed in the process using known mathematical relations. Calorimeters are designed to minimize energy exchange between the system being studied and its surroundings. They range from simple coffee cup calorimeters used by introductory chemistry students, to sophisticated bomb calorimeters used to determine the energy content of food.

8.3: Enthalpy

If a chemical change is carried out at constant pressure and the only work done is caused by expansion or contraction, **q** for the change is called the **enthalpy change** with the symbol **ΔH** , or **ΔH°_{298}** for reactions occurring under standard state conditions. The value of ΔH for a reaction in one direction is equal in magnitude, but opposite in sign, to ΔH for the reaction in the opposite direction, and ΔH is directly proportional to the quantity of reactants and products. Examples of enthalpy changes include enthalpy of combustion, enthalpy of fusion, enthalpy of vaporization, and standard enthalpy of formation. The **standard enthalpy of formation, ΔH_f°** , is the enthalpy change accompanying the formation of 1 mole of a substance from the elements in their most stable states at 1 bar (standard state). Many of the processes are carried out at 298.15 K. If the enthalpies of formation are available for the reactants and products of a reaction, the enthalpy change can be calculated using **Hess's law**: If a process can be written as the sum of several stepwise processes, the enthalpy change of the total process equals the sum of the enthalpy changes of the various steps.