Thermodynamics
Chem1B, General Chemistry II

- **System:** What you are studying (usually the reaction)
- **Surroundings:** Everything else (outside solution, beaker, air, etc.)
- **Universe** = system + surroundings

The **system** and **surroundings** always work opposite each other.

      If the **system** gains, the **surroundings** lose

      If the **system** loses, the **surroundings** gain

State Functions are path-independent and always capitalized (i.e., elevation)

Path-dependent functions change based on distance and are lower case (i.e., work done)

- **Energy (E):** the ability to do work
- **Kinetic Energy (KE):** related to movement or motion
- **Potential Energy (PE):** related to position or composition
- **Total Energy** = kinetic energy (KE) + potential energy (PE)

Internal Energy (ΔE/U): sum of kinetic and potential energy in a **system**

      Always measured as the *change* ($\Delta E = E_{final} - E_{initial}$)

In a system, $\Delta E = q + w$

- **Heat (q)**: transfer of thermal energy
- **Work (w)**: energy used to move an object against a force

For a system, the signs matter.

      If q is **absorbed**, the sign is **positive**. If q is **released,** the sign is **negative**.

      If w is done **on** a system, the sign is **positive.** If w is done **by** a system, the sign is **negative.**

For gases, work is specifically **pressure-volume (PV) work**, $w = -P\Delta V$

      If a gas **expands** (+ΔV) against an external pressure, the system is **doing** work.

      If a gas **compresses** (-ΔV) by an external pressure, the system is **being worked on**.

Enthalpy (ΔH): heat flow in a constant-pressure (**isobaric**) system

      $\Delta H = H_{products} - H_{reactants}$

In a system, $\Delta H = \Delta E + P\Delta V$

      With constant pressure, $\Delta E = q + w$ and $w = -P\Delta V$

      so that $\Delta H = (q + w) + (- w) = q_p$

The sign on q determines the sign on ΔH.

      If the system absorbs heat, ΔH is **positive (endothermic).**

      If the system releases heat, ΔH is **negative (exothermic)**.

Four ways to determine the enthalpy of a reaction:

1.) Calorimetry (coffee-cup or bomb, using $q = mC_s\Delta T$)
2.) Hess's Law (known heats of reactions summed together)
3.) Enthalpies of Formation ($\Delta H°_{rxn} = \sum\Delta H°_{products} - \sum\Delta H°_{reactants}$)
4.) Bond Enthalpies ($\Delta H_{rxn} = \sum\Delta H_{bonds\ broken} - \sum\Delta H_{bonds\ formed}$)

Standard Conditions (°): P = 1 atm, T = 298 K, concentrations = 1 M