Thermodynamics

Lecture Notes: Internal Energy

Definition of Internal Energy

- **Internal Energy (U):** The total energy contained within a system.
  - It includes all forms of kinetic and potential energies of particles (atoms, ions, molecules) in the system.
  - Examples: Molecular vibrations, bond energies, translational and rotational energies.

Significance of Internal Energy

1. **State Function**:
   - Internal energy is a *state function*, meaning its value depends only on the current state of the system (temperature, pressure, volume) and not on how that state was achieved.
   - It is represented as U, and changes in internal energy (ΔU) provide valuable insights into system transformations.
2. **Key Role in Thermodynamics**:
   - Internal energy is central to the **First Law of Thermodynamics**, which relates changes in energy to heat (q) and work (w):

   $$\Delta U = q + w$$

   - This equation underlines energy conservation within a system.
3. **Determining System Behavior**:
   - The value of U helps in understanding:
     - Energy transfers during chemical reactions.
     - Energy changes in physical processes like phase transitions.
4. **Thermal Equilibrium**:
   - Systems at the same temperature have equivalent internal energy under identical conditions.

Work and Heat Exchange

1. **Work (w):**
   - Work is the energy transfer when a force is applied over a distance.
   - In thermodynamic systems:
     - **Expansion/Compression Work:** Work done by or on the system when its volume changes.

     $$w = -P\Delta V$$

     - P: External pressure, ΔV: Volume change.

- Example: Gas expansion in a piston.

2. **Heat (q):**
    - Heat is the energy transfer due to temperature differences between the system and surroundings.
    - Heat exchange modifies the system's internal energy:
        - **Heat Absorbed ($+q$):** Increases U.
        - **Heat Released ($-q$):** Decreases U.

Relationship Between Work and Heat

- Both work and heat are mechanisms of energy transfer into or out of a system.
- The **First Law of Thermodynamics** combines them:

$$\Delta U = q + w$$

    - $q > 0$: Heat absorbed by the system.
    - $w > 0$: Work done on the system.
    - $q < 0$: Heat released from the system.
    - $w < 0$: Work done by the system.

Applications

- **Chemical Reactions**:
    - Calculating heat changes (q) helps determine reaction enthalpies.
- **Physical Processes**:
    - Understanding work and heat exchange is critical for processes like refrigeration and engine cycles.
- **Energy Conservation**:
    - Ensures no energy is created or lost, only transformed or transferred.
