Introduction to Thermodynamics

Chapter 2 of Atkins: The First Law: Concepts

Sections 2.1-2.2 of Atkins

Basic Concepts

- Work
- Heat
- Energy

The First Law of Thermodynamics - Conservation of Energy
**Introduction to Thermodynamics**

**Thermodynamics**: Study of static (i.e., stationary) states of matter and how they differ from one another in terms of energetic quantities.

Thermodynamics gives us a set of rules which applies to any object or set of objects independent of what those objects are, what they were previously, or what they will be in the future.

There are some (relatively) simple assumptions and definitions, and a mathematical framework available, which allows scientists and engineers to take very complex situations and reduce them to relatively simple problems.

**Applications:**
- Heating and cooling buildings
- Batteries (chemical work)
- Plant/ factory design
- Engine efficiency
- Biological energy transfer
- Insulators and conductors

Food

Energy

Waste
Thermodynamics: The study of patterns of energy change, where *thermo* refers to heat, and *dynamics* refers to patterns of change

(a) energy conversion
(b) directions of change and molecular stability

(a) $U_p = mgh$

(b) *Protein structure and folding*

*heat, sound, light* upon impact
**Systems and Surroundings**

**System**: The object or region of interest

**Surroundings**: Everything outside of the system (we do our measuring in the surroundings)

The definition of the system is very dependent upon the **boundaries** which separate the system from the surroundings - i.e., can **energy**, which is the capacity to do work, be transferred between the system and surroundings.

**Open**: Matter can go in and out of the system

**Closed**: Matter cannot go in and out of the system

**Isolated System**: Closed system with no thermal or mechanical contact with the outside world (this is an **ideal system**)
**Work, Heat and Energy**

**Energy:** Capacity to do work - the energy of a system can be changed by work and heat  
**Units:** joules (J) or for molar energy kJ mol\(^{-1}\)

**Work:** A form of energy which can transfer in and out of a system, that is stored in the *organized motion* of molecules. Work is done when an object is moved against some opposing force.

**Heat:** A form of energy which can transfer in and out of a system, that is stored in the *random motion* (thermal motion) of molecules. If an energy difference of a system results from a temperature difference between it and its surroundings, energy has been transferred as heat.

---

Prof. Mueller  
Chemistry 451 - Fall 2003  
Lecture 5 - 5
Thermodynamic Definitions of Work & Heat

Work: Work \((w)\) is energy which is transferred across the boundary of a thermodynamic system during a change in its state and which could be converted completely into the “lifting of a weight” in the surroundings (i.e., motion of a mass against opposing force).

Sign conventions for work:

\[
\begin{align*}
\text{System} & \quad w_{\text{system}} < 0 \\
\text{System} & \quad w_{\text{system}} > 0
\end{align*}
\]

Heat: Heat \((q)\) is energy which is transferred across the boundary of a thermodynamic system during a change in its state by virtue of a difference in temperature between the system and its surroundings.

Sign conventions for heat:

\[
\begin{align*}
\text{System} & \quad q_{\text{system}} < 0 \\
\text{System} & \quad q_{\text{system}} > 0
\end{align*}
\]
More About Heat

**Diathermic (Diabatic) System**: A system in which energy is allowed in and out in the form of heat - system in **thermal contact** with surroundings

**Adiabatic System**: A system in which heat is not allowed in or out

**Endothermic Process**: (a)
Process that absorbs heat

**Exothermic Process**: (b)
Process that releases heat

Endothermic process in a diathermic container (c): decreases temperature of surroundings

Exothermic process in a diathermic container (d): increases temperature of surroundings
Internal Energy

**Internal Energy** \((U)\): total energy of a *system* at any given time – it comes from the total kinetic and potential energy of molecules which compose the system.

**Change in Energy** \((\Delta U)\): Change in \(U\) as system goes from an initial state with energy \(U_i\) to final state with \(U_f\).

\[
\Delta U = U_f - U_i
\]

*Internal energy is:*

- **a state function**: only depends on the current state, not on preparation of the state.

- **an extensive property**: is dependent upon the amount of matter. Properties whose values are increased/decreased in direct proportion to the enlargement/reduction of the system.

Remember, an **intensive property** is independent of the amount of matter. All the thermodynamic driving forces are selected from among properties of this type.
Molecular Interpretation of Energy

Energy of one atom of a monatomic perfect gas at temperature T:

According to Boltzmann, average energy of each term is \( \frac{1}{2} kT \); mean energy of atoms is three times this, with total energy \( \frac{3}{2} NkT = \frac{3}{2} nRT \):

here \( U_m(0) \) is the internal molar energy at \( T=0 \) (no translation) - so energy can also come from internal structure of atoms.

If a gas has polyatomic molecules, they can rotate about three axes as well, for an additional energy contribution of \( \frac{3}{2} RT \) (the energy increases twice as rapidly as above).

**Bottom line is:** as \( T \) is increased the internal energy increases as various modes of motion are excited.
Conservation of Energy

Internal energy of a system can be changed by:
1. Doing work to the system (w)
2. Energy transferred as heat to the system (q)

Heat and work are equivalent ways of changing the energy of a system

System is like a bank, where currencies can be deposited or withdrawn

If a system is completely isolated from its surroundings: no change in the internal energy takes place: $\Delta U=0$ (no perpetual motion)

Recall:

If heat or work are transferred to the system:
\[ q>0, \quad w>0 \] (system acquires energy)

If heat or work are transferred from the system:
\[ q<0, \quad w<0 \] (system releases energy)
James Prescott Joule is the scientist credited with establishing the First Law of Thermodynamics. He was born on December 24, 1818 in Salford, England. As a child, Joule was educated at home, and he never received a formal education. When Joule was fourteen, his father began sending him to weekly chemistry lessons with the famous scientist John Dalton. By 1838, Joule had begun performing his own experiments, and his first paper was published that year. Throughout his life, Joule performed many experiments in the area of thermodynamics, and his determination of the principle of conservation of energy (essentially the First Law of Thermodynamics) is considered to be the chief contribution of the nineteenth century to physical science.

In the past, there was some belief in conservation of matter (except for combustion), and Lavoisier demonstrated this in chemical reactions.
First Law of Thermodynamics

The equation $\Delta U = q + w$ is a mathematical statement of the First Law of Thermodynamics, stating that

(i) Heat and work are equivalent forms of energy, and are the only ways to change the internal energy of a system

(ii) Internal energy is conserved in any thermodynamic process in an isolated system, $\Delta U = 0$, since $q$ and $w$ must both be zero

*Alternative Statement of the First Law of Thermodynamics:*

In any thermodynamic process, the internal energy of the universe is conserved, as energy is neither created nor destroyed:

This follows from the law above, since the universe can be considered an isolated system.
Energy is a property of any system, and changes in this property can be deduced from the conservation-of-energy principle. The work needed to change an adiabatic system from one specified state to another is the same however the work is done (mountain climbing analogy).

So, internal energy can be measured by measuring work needed to bring about change in an adiabatic system.

If 1 kJ (kilojoule = 1000 J) of work is done by stirring a system, or 1 kJ of electrical work is done by passing current through a heater, the same increase in energy (and therefore rise in temperature) is produced.
Mechanical Definition of Heat

Imagine that our system is not purely adiabatic (thermal insulation is not very good). The change in $U$, the internal energy, may be the same as before (state function), but more work is required to make the same kind of change in energy:

42 kJ of work in an adiabatic container;
however, to achieve the same change of state in a diathermic container: 50 kJ of work is required.

The system is absorbing heat if $q > 0$

\[
q = w_{ad} - w
\]
\[
q = 42 \text{ kJ} - 50 \text{ kJ} = -8 \text{ kJ} \quad (8 \text{ kJ is lost from the system as heat})
\]

Since $\Delta U$ is by definition equal to $w_{ad}$, the expression for energy transferred to a system as heat is

\[
q = \Delta U - w
\]

which is the same as the expression for the First Law, $\Delta U = q + w$.
The *First Law* works differently for a process in which the volume stays constant than for a process in which the pressure stays constant.

**Isochoric Process**
The volume of the system remains constant ($\Delta V=0$). On a $pV$ diagram, this type of process would be represented by a vertical line. Therefore, there is no area under the curve.

Since there is no change in volume, $w = 0$. So the internal energy of the system depends only on heat transfer.

**Isobaric Process**
The system's pressure is fixed ($\Delta P=0$). It is represented by a horizontal line on a $pV$ diagram.

Both $w$ and $q$ will still exist for this process.
Indicator Diagrams

Used to determine expansion work

![Diagram showing isobar and isotherm](image-url)
Applications/Processes of Thermodynamics

Adiabatic Process
There is no transfer of heat to or from the system, \( q = 0 \), though other characteristics such as the pressure, volume, and temperature of the system may vary. The curve of an adiabatic process on a \( pV \) diagram depends on the changes in the system.

Isothermal Process
The system’s temperature is fixed (\( \Delta T = 0 \)). It is represented by a \( pV \) diagram as we have seen earlier and energy stays constant for an ideal gas (\( \Delta U = 0 \))

Cyclic Process
Certain processes have the same initial and final states (e.g., engines). This is represented by a closed path and \( \Delta U = 0 \)