To understand and perform any sort of thermodynamic calculation, we must first understand the fundamental laws and concepts of thermodynamics. For example, work and heat are interrelated concepts. Heat is the transfer of thermal energy between two bodies that are at different temperatures and is not equal to thermal energy. Work is the force used to transfer energy between a system and its surroundings and is needed to create heat and the transfer of thermal energy. Both work and heat together allow systems to exchange energy. The relationship between the two concepts can be analyzed through the topic of Thermodynamics, which is the scientific study of the interaction of heat and other types of energy.

Introduction

To understand the relationship between work and heat, we need to understand a third, linking factor: the change in internal energy. Energy cannot be created nor destroyed, but it can be converted or transferred. Internal energy refers to all the energy within a given system, including the kinetic energy of molecules and the energy stored in all of the chemical bonds between molecules. With the interactions of heat, work and internal energy, there are energy transfers and conversions every time a change is made upon a system. However, no net energy is created or lost during these transfers.

Definition: 1st Law of Thermodynamics

The First Law of Thermodynamics states that energy can be converted from one form to another with the interaction of heat, work and internal energy, but it cannot be created nor destroyed, under any circumstances. Mathematically, this is represented as

\[ ΔU = q + w \]

with

- \( ΔU \) is the total change in internal energy of a system,
- \( q \) is the heat exchanged between a system and its surroundings, and
- \( w \) is the work done by or on the system.

Work is also equal to the negative external pressure on the system multiplied by the change in volume:

\[ w = -pΔV \]

where \( p \) is the external pressure on the system, and \( ΔV \) is the change in volume. This is specifically called "pressure-volume" work.

The internal energy of a system would decrease if the system gives off heat or does work. Therefore, internal energy of a system increases when the heat increases (this would be done by adding heat into a system). The internal energy would also increase if work were done onto a system. Any work or heat that goes into or out of a system changes the internal energy. However, since energy is never created nor destroyed (thus, the first law of thermodynamics), the change in internal energy always equals zero. If energy is lost by the system, then it is absorbed by the surroundings. If energy is absorbed into a system, then that energy was released by the surroundings.
\[
\Delta U_{\text{system}} = -\Delta U_{\text{surroundings}}
\]

where \(\Delta U_{\text{system}}\) is the total internal energy in a system, and \(\Delta U_{\text{surroundings}}\) is the total energy of the surroundings.

Table 1

<table>
<thead>
<tr>
<th>Process</th>
<th>Sign of heat (q)</th>
<th>Sign of Work (w)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Work done by the system</td>
<td>N/A</td>
<td>-</td>
</tr>
<tr>
<td>Work done onto the system</td>
<td>N/A</td>
<td>+</td>
</tr>
<tr>
<td>Heat released from the system- exothermic</td>
<td>-</td>
<td>N/A</td>
</tr>
<tr>
<td>(absorbed by surroundings)</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

The above figure is a visual example of the First Law of Thermodynamics. The blue cubes represent the system and the yellow circles represent the surroundings around the system. If energy is lost by the cube system then it is gained by the surroundings. Energy is never created nor destroyed. Since the area of the cube decreased the visual area of the yellow circle increased. This symbolizes how energy lost by a system is gained by the surroundings. The affects of different surroundings and changes on a system help determine the increase or decrease of internal energy, heat and work.

Table 2

<table>
<thead>
<tr>
<th>The Process</th>
<th>Internal Energy Change</th>
<th>Heat Transfer of Thermal Energy (q)</th>
<th>Work (w=-P\Delta V)</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>(q=0) Adiabatic</td>
<td>+</td>
<td>0</td>
<td>+</td>
<td>Isolated system in which heat does not enter or leave similar to styrofoam</td>
</tr>
<tr>
<td>(\Delta V=0) Constant Volume</td>
<td>+</td>
<td>+</td>
<td>0</td>
<td>A hard, pressure isolated system like a bomb calorimeter</td>
</tr>
<tr>
<td>Constant Pressure</td>
<td>+ or -</td>
<td>enthalpy ((\Delta H))</td>
<td>-(P\Delta V)</td>
<td>Most processes occur are constant external pressure</td>
</tr>
<tr>
<td>(\Delta T=0) Isothermal</td>
<td>0</td>
<td>+</td>
<td>-</td>
<td>There is no change in temperature like in a temperature bath</td>
</tr>
</tbody>
</table>

Example \(\PageIndex{1}\)

A gas in a system has constant pressure. The surroundings around the system lose 62 J of heat and does 474 J of work onto the system. What is the internal energy of the system?

**SOLUTION**

To find internal energy, \(\Delta U\), we must consider the relationship between the system and the surroundings. Since the First Law of Thermodynamics states that energy is not created nor destroyed we know that anything lost by the surroundings
is gained by the system. The surrounding area loses heat and does work onto the system. Therefore, \(q\) and \(w\) are positive in the equation \(\Delta U = q + w\) because the system gains heat and gets work done on itself.

\[
\begin{align}
\Delta U &= (62\,J) + (474\,J) \\
&= 536\,J 
\end{align}
\]

Example \(\PageIndex{2}\))

A system has constant volume (\(\Delta V = 0\)) and the heat around the system increases by 45 J.

a. What is the sign for heat (\(q\)) for the system?
b. What is \(\Delta U\) equal to?
c. What is the value of internal energy of the system in Joules?

**SOLUTION**

Since the system has constant volume (\(\Delta V = 0\)) the term \(-P\Delta V = 0\) and work is equal to zero. Thus, in the equation \(\Delta U = q + w\) \(w = 0\) and \(\Delta U = q\). The internal energy is equal to the heat of the system. The surrounding heat increases, so the heat of the system decreases because heat is not created nor destroyed. Therefore, heat is taken away from the system making it exothermic and negative. The value of Internal Energy will be the negative value of the heat absorbed by the surroundings.

a. negative \((q < 0)\)
b. \(\Delta U = q + (-P\Delta V) = q + 0 = q\)
c. \(\Delta U = -45J\)

**Outside Links**


**References**


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