First Law of Thermodynamics

The first law of thermodynamics centres on the concept of energy. In its broadest sense, the law requires that the energy of the universe is constant. This is a rather overwhelming statement. A more attractive statement requires that the (internal) thermodynamic energy $U$ of a chemistry laboratory is constant:

$$U = \text{constant} \quad (a)$$

The latter statement is the **principle of conservation of energy**: energy can be neither created nor destroyed. All that a chemist can do, during an experiment using a closed reaction vessel, is to watch energy ‘move’ between system and surroundings. As a consequence of equation (a), we state that

$$\Delta U_{\text{system}} = -\Delta U_{\text{surroundings}} \quad (b)$$

We can not know the actual energy of a closed system although we know that it is an extensive property of the system [1]. In describing energy changes we need a convention. So we use the **acquisitive convention** describing all changes in terms of how a system is affected. Thus the statement $\Delta U < 0$ means that the energy of a given system decreases during a given process; e.g. chemical reaction.

**Footnote**

[1] In principle it is possible to know the total energy of a given system using a scale in conjunction with Einstein’s famous equation, $E = m \cdot c^2$. However, the mass corresponding to 1 kJ is only about $10^{-14}$ kg.