

Chapter 18: The First Law of Thermodynamics

Temperature measures the amount of energy of the particles of a given an object, which we feel as hot with more energy, and cold with less energy. We measure temperature in Kelvin (K), but also use the celsius scale ($^{\circ}\text{C}$), and the fahrenheit scale ($^{\circ}\text{F}$):

$$T_C = T_K - 273.15^{\circ}\text{C} \quad (\text{Celsius})$$

$$T_F = \frac{9}{5}T_C + 32^{\circ}\text{C} \quad (\text{Fahrenheit})$$

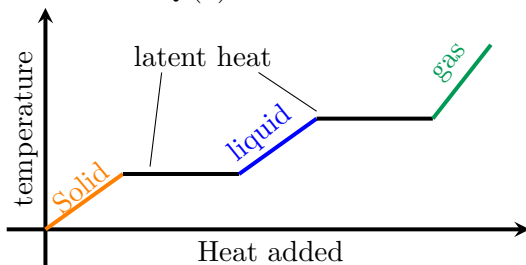
Temperature affects the physcal properties of an object, such as heat causing metal to expand. We can calculate the change in length over a change in temperature ($^{\circ}\text{C}$ or K):

$$\Delta L = L\alpha\Delta T \quad (\text{linear expansion})$$

Where α and $\beta = 3\alpha$ is the coefficient of linear expansion ($^{\circ}\text{C}^{-1}$). For a change in volume we use:

$$\Delta V = V\beta\Delta T \quad (\text{volume expansion})$$

Now if we have an object in an environment, such as an iced soda can outside in the sun, its temperature rises quickly at first, but then the soda's temperature increasing rate slows down until it reaches the ambient temperature. We see that if the temperature of our system T_s (the soda) and the environment T_e are different, they will reach equilibrium. This change of temperature is the result of energy transfer called Heat Q (J).



When heat is added to system, the temperature of an object *tends* to increase until an object reaches melting or vaporizing point where it will need more energy to change states. The latent heat L (cal/g or J/kg) of a material is the amount of energy required to change states:

$$Q = mL \quad (\text{latent heat})$$

Where the **latent heat of fusion** L_f is the transition between **solid and liquid**, while the **latent heat of vaporization** L_v is the transition between **gas and liquid**. Aside from heat, Calories and British thermal units (Btu) measure the amount of energy that would raise the temperature of 1g of water from 14.5°C to 15.5°C and 1 lb of water from 63°F to 64°F respectively.

$$1 \text{ cal} = 3.968 \times 10^{-3} \text{ Btu} = 4.1868 \text{ J}$$

With this, we can find the amount of heat used during a change of temperature, given the heat capacity C (J/K) of an object.

$$Q = C\Delta T = C(T_f - T_i) \quad (\text{heat capacity})$$

However, the heat capacity is different for every system, depending on the material of an object, and its mass. The heat capacity is the product of the heat specific heat c ($\frac{\text{J}}{\text{kg}\cdot\text{K}}$) and its mass. With this, the specific heat unit conversions are now:

$$1 \text{ cal/g} = 3.968 \times 10^{-3} \text{ Btu/lb} = 4.1868 \text{ J}/(\text{K} \times \text{kg})$$

Now we can find heat used over a change of temperature given the specific heat constant and mass:

$$Q = cm\Delta T = cm(T_f - T_i) \quad (\text{specific heat})$$

Remembering that heat is energy, we can use Power ($W = \frac{J}{s}$) and time:

$$Q = Pt \quad (\text{power})$$

Chapter 18: The First Law of Thermodynamics

As energy can neither be created nor destroyed, the first Law of Thermodynamics states:

$$\Delta E_{int} = E_f - E_i = Q - W \quad (\text{first law})$$

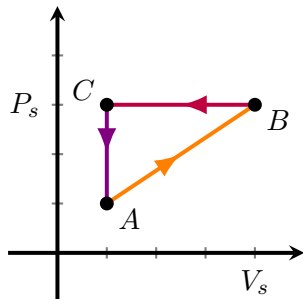
And if the system undergoes differential change:

$$dE = dQ - dW \quad (\text{first law})$$

The principle of energy conservation shows that the change of our internal energy in the system (ΔE_{int}) is found by heat added to system Q while removing the work done by the system. Imagine you are charging your phone, causing energy to transfer to the phone; this is work done on your phone (system). Now if your phone was playing a movie, it expends energy, causing work by your phone (system). With this in mind, we can start working with how gas can exchange energy with its surroundings through work:

$$W = \int dW = \int_{v_i}^{v_f} p \, dV \quad (\text{work})$$

In a system such as an engine, a piston will compress gas in the engine, doing work on the system by causing a change in volume. Since work is a force done over a displacement, $dW = F \cdot dx$, we can rewrite the force in terms of area and pressure $F = pA$. This gives us $dW = pA \cdot dx = p \, dV$



Here we have a graph of a $p - V$ diagram where a gas undergoes compression and expansion. Let $P_s = 30 \text{ N/m}^2$ and $V_s = 4.0 \text{ m}^3$. To find the net energy added to the system, we need to find how much work was done for each segment, W_{AB} , W_{BC} and W_{CA} . Lets start with W_{BC} , since pressure p is constant through the path:

$$W_{BC} = \int_{V_s}^{\frac{1}{4}V_s} p \, dV = P_s W \Big|_{V_s}^{\frac{1}{4}V_s} = P_s \left[\frac{1}{4}V_s - V_s \right] = P_s \left[-\frac{3}{4}V_s \right] = -90 \text{ J}$$

We see that the system did 90J of work. Now let's take a look at W_{CA} . Note that the volume does not change, so the integral's bounds will result in 0, giving $W_{CA} = 0$. Lastly, to find W_{AB} we will need to treat pressure as a function of volume. In this case we have two points A and B , so we will use the line formula to give us this integral:

$$W_{AB} = \int_{\frac{1}{4}V_s}^{V_s} \frac{10}{3} + \frac{20}{3}V \, dV = 60 \text{ J}$$

Now the net work $W_{net} = W_{AB} + W_{BC} + W_{CA}$ which gives us -30 J . Given that $\Delta E_{int} = -W = 30 \text{ J}$, this means that 30J of energy was done on the system, increasing its heat. In this case, we did not consider the transfer of heat between the environment and the system, resulting in $Q = 0$, know as an adiabatic process. In fact 4 special cases for the first law of Thermodynamics:

$$\Delta E_{int} = -W \quad (\text{adiabatic process})$$

When no transfer of heat between the environment and the system.

$$\Delta E_{int} = Q \quad (\text{Constant-volume process})$$

When the volume of a system is constant resulting in $W = 0$.

$$\Delta Q = W \quad (\text{Cyclical process})$$

After a process has occurred, returning to its original state, resulting in the system's internal energy.

$$\Delta E_{int} = 0 \quad (\text{Free expansions})$$

Chapter 18: The First Law of Thermodynamics

If we leave the iced soda can outside, the energy is transferred from the outside environment to the inside through **conduction** (W):

$$P_{cond} = \frac{Q}{t} = kA \cdot \frac{\Delta T}{L} \quad (\text{conduction})$$

Where k is the conductivity of your material, A is the area of the barrier, and L is the length. As with conduction we can also measure the resistance, or how good a material at a specific length is at insulating heat:

$$R = \frac{L}{K} \quad (\text{resistance})$$

Sometimes there will be multiple layers of a different materials between hot and cold areas. We can use the resistance of each material to measure the whole energy transfer rate:

$$P_{cond} = \frac{A\Delta T}{\sum \frac{L}{K}} \quad (\text{conduction})$$

If we have a candle in a cold environment, we can feel the heat radiating out of the candle. This energy transfer is **thermal radiation** which is when energy is exchanged through electromagnetic waves. The rate at the energy is emitted by the system is given by:

$$P_{rad} = \alpha \epsilon A T_{obj}^4 \quad (\text{radiation rate})$$

Here, the temperature of the object is in Kelvin, $\alpha = 5.6704 \times 10^8 \text{W/m}^2 \cdot \text{K}^4$ is the Stefan–Boltzmann constant and ϵ is the emissivity of the object's surface. However, an object can also absorb thermal radiation:

$$P_{abs} = \alpha \epsilon A T_{env}^4 \quad (\text{absorption rate})$$

With this we can find the net power transfer where a positive value signifies energy is absorbed while a negative value shows that energy is lost:

$$P_{net} = P_{abs} - P_{rad} = \alpha \epsilon A (T_{env}^4 - T_{obj}^4) \quad (\text{net power})$$