Thermodynamics

Josh Borrow

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0.1 Introduction

Thermodynamics predates the atomic model and was created so that we can get the most work out of machines for a given input. We began by studying bulk material properties (think of a black box), however this is generalistic and it's difficult to see the point. We can know a heat capacity, great. This doesn't tell us anything about the physics that's going on, though.

In this course we will open the black box and look at the microscopic picture. This really enriches our understanding, but we have to use statistical mechanics. In a lecture hall, there are about 10^{30} atoms. With a 5Ghz processor, this would take over the age of the universe to count (one per clock cycle!) So what we really study is the average behaviour of many-bodied systems.

0.2 Definitions

Here we will use two systems to compare *extensive* and *intensive* variables.

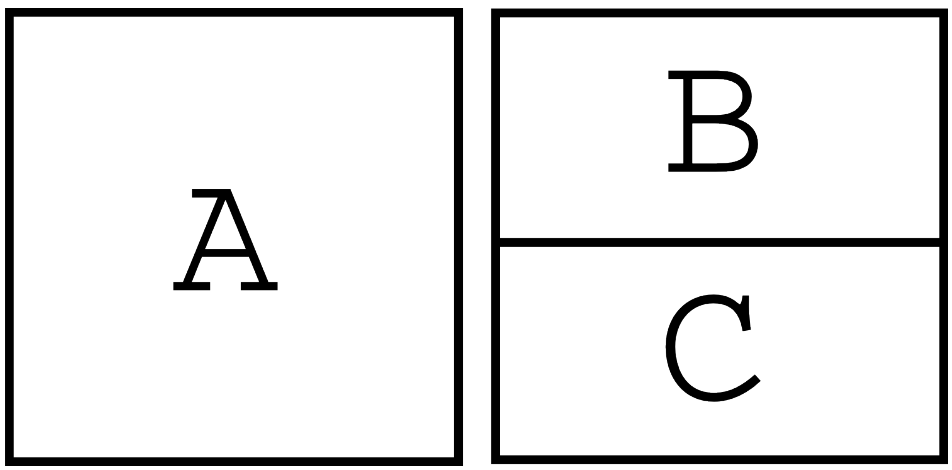


Figure 1: Three systems that start all with the same initial conditions

Extensive variables are ones which change when the system is split up. For example:

- Volume $V_A = 2V_B = 2V_C$
- Energy $U_A = 2U_B = 2U_C$

Intensive variables are independent of the size of the system:

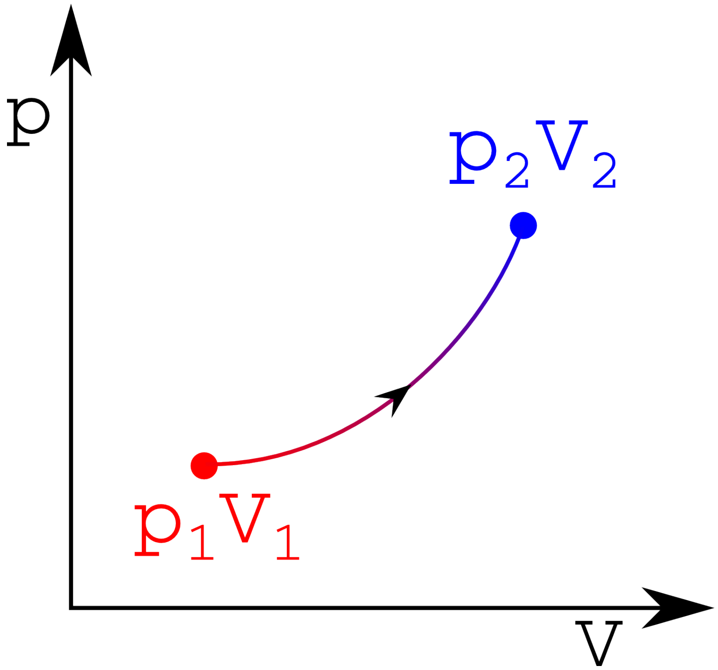
- Temperature $T_A = T_B = T_C$
- Pressure $P_A = P_B = P_C$

The system is the matter we're interested in.

The surroundings is the rest of the universe.

0.2.1 P-V Diagrams

P-V diagrams are favourites in the thermodynamics community. They show the path between two points of known pressure and volume:



Temperature is a measure of how hot something is.

Heat is energy in transit.

0.3 Equations of state

Equations of state describe a system mathematically:

$$f = f(p, V, T, \ldots)$$

For example:

$$pV = RT \;\rightarrow\; f = f(p, V)$$

Here we have two independent variables, and we can always find the third.

0.4 Heat

Heat always travels from a hot body to a cold body, if there is no external influence.

0.4.1 Heat capacity

Heat capacity is a measure of how much energy is needed to increase the temperature of a system:

$$C_v = \left(\frac{\partial Q}{\partial T}\right)_V$$

We can then find the change in energy needed by integrating:

$$\Delta Q = \int_{T_1}^{T_2} C_v \cdot \mathrm{d}T$$

There are two types of heat capacity that we use regularly. There is the heat capacity at constant volume (C_V), and the heat capacity at constant pressure (C_p). The heat capacity at constant pressure is always higher because we need to do work in increasing the volume of the gas, as well as the temperature.

0.5 The Zeroth Law of Thermodynamics

This law is the basis of temperature. It says:

If we have three systems, A, B and C, and we begin by letting A and B thermally equilibrate, then we have two systems in thermal equilibrium. If we then move B to C and there is no net heat transfer, meaning that B and C are already in equilibrium, then this means that A and C are in thermal equilibrium too - meaning they must be the same *temperature*.