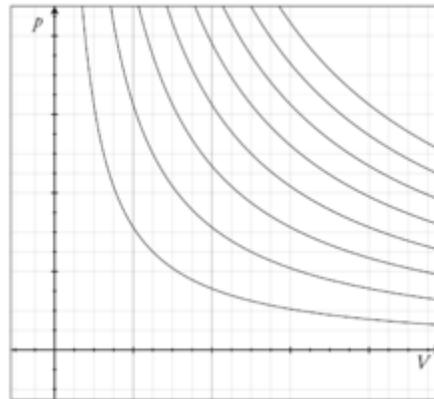


Quantum Mechanics_isothermal process

An **isothermal process** is a change of a *system*, in which the temperature remains constant: $\Delta T = 0$. This typically occurs when a system is in contact with an outside thermal reservoir (heat bath), and the change occurs slowly enough to allow the system to continually adjust to the temperature of the reservoir through Heat exchange. In contrast, an adiabatic process is where a system exchanges no heat with its surroundings ($Q = 0$). In other words, in an isothermal process, the value $\Delta T = 0$ and therefore $\Delta U = 0$ but $Q \neq 0$, while in an adiabatic process, $\Delta T \neq 0$ but $Q = 0$.

Details for an ideal gas



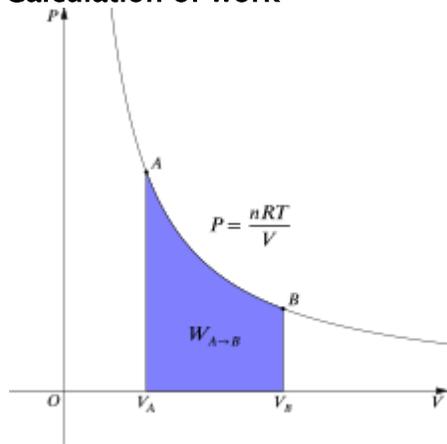
Several isotherms of an ideal gas on a p-V diagram

For the special case of a gas to which Boyle's law applies, the product pV is a constant if the gas is kept at isothermal conditions. However, in the cases where the product pV is an exponential term this does not comply. The value of the constant is nRT , where n is the number of moles of gas present and R is the ideal gas constant. In other words, the Ideal gas law $pV = nRT$ applies. This means that

$$p = \frac{nRT}{V} = \frac{\text{constant}}{V}$$

holds. The family of curves generated by this equation is shown in the graph presented at the bottom right-hand of the page. Each curve is called an isotherm. Such graphs are termed indicator diagrams and were first used by James Watt and others to monitor the efficiency of engines. The temperature corresponding to each curve in the figure increases from the lower left to the upper right.

Calculation of work



The purple area represents "work" for this isothermal change

In thermodynamics, the Work involved when a gas changes from state *A* to state *B* is simply

$$W_{A \rightarrow B} = - \int_{V_A}^{V_B} p dV$$

For an isothermal, reversible process, this integral equals the area under the relevant pressure–volume isotherm, and is indicated in purple in the figure (at the bottom right-hand of the page) for an ideal gas. Again, $p = nRT / V$ applies and with T being constant (as this is an isothermal process), we have:

$$W_{A \rightarrow B} = - \int_{V_A}^{V_B} p dV = - \int_{V_A}^{V_B} \frac{nRT}{V} dV = -nRT \int_{V_A}^{V_B} \frac{1}{V} dV = -nRT \ln \frac{V_B}{V_A}$$

By convention, work is defined as the work the system does on its environment. If, for example, the system expands by a piston moving in the direction of force applied by the internal pressure of a gas, then the work is counted as positive, and as this work is done by using internal energy of the system, the result is that the internal energy decreases. Conversely, if the environment does work on the system so that its internal energy increases, the work is counted as negative.

It is also worth noting that, for many systems, if the temperature is held constant, the internal energy of the system also is constant, and so $\Delta U = 0$. From First Law of Thermodynamics, $\Delta U = Q - W$, so it follows that $Q = W$ for this same isothermal process.

When no heat flows into or out of the gas because its container is at the same temperature, then there is no work done. Thus, work=0 which means external pressure on any moving surface is zero. This is called Free expansion.

Applications

Isothermal processes can occur in any kind of system, including highly structured machines, and even living cells. Various parts of the cycles of some Heat engines are carried out isothermally and may be approximated by a Carnot cycle. Phase changes, such as melting or evaporation, are also isothermal processes.

In Isothermal non flow Process, the work done by compressing the perfect gas (Pure Substance) is a negative work, as work is done on the system, as result of compression, the volume will decrease, and temperature will try to increase. To maintain the temperature at constant value (as the process is isothermal) heat energy has to leave the system and enter the environment. The amount of energy entering the environment is equal to the work done (by compressing the perfect gas) because internal energy does not change. The thermodynamic sign convention is that heat entering the environment is also negative. Thence $-Q = W$.

In equation of work, the term nRT can be replaced by PV of any state of an ideal gas. The product of pressure and volume is in fact, '**Moving Boundary Work**'; the systems boundaries are compressed. For Expansion the same theory is applied.

As per Joule's Law for the perfect gas, Internal energy is the function of absolute temperature. In an Isothermal process the temperature is constant. Hence, the internal energy is constant, and the net change in internal energy is ZERO. Within the perfect, or ideal gas, there are no inter-molecular forces and the gas particles are infinitesimal. However, for a real pure substance there is a component of internal energy corresponding to the energy used in overcoming inter-molecular forces. In an Isothermal process, when the volume of the gas changes, the average distance between each molecule changes as well. So if the real pure gas undergoes an Isothermal process, there is a net change in internal temperature consistent with this component of internal energy.[1]

Notes

1. Δ Adkins 1983, p. 121.

References

- Adkins, C. J.(1983). *Equilibrium Thermodynamics*. Cambridge University Press.

Source: <http://waterkalinemachine.com/quantum-mechanics/?wiki-maping=Isothermal>