Complex-Information Theory

Bridging Physics to Ethics

By

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This book first published 2025

Ethics International Press Ltd, UK

British Library Cataloguing in Publication Data

A catalogue record for this book is available from the British Library

Print Book ISBN: 978-1-83711-049-0

eBook ISBN: 978-1-83711-050-6

For Deborah

Sincere gratitude to David LePoire, Kevin Gibson, and Todd Duncan for their reviews, comments, and corrections.

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Preface

An Intersection between Philosophy and Science

I understood before writing this book, that proposing a viable and fruitful intersection of disciplines stretching from ethics “down to” a basic science like physics might be provocative to potential readers. My only personal glimpse of this possibility was a retort I had gotten from a philosopher acquaintance after giving her a brief description of “complex-information theory”: “Hah! Fatal flaw, physics is dead!” However, after reading the Afterword of Sam Harris’ book, *The Moral Landscape: How Science Can Determine Human Values* (Harris, S., 2009), I had a better realization that I could be “poking a hornet’s nest” more than I surmised.

I was unaware of his book until I had nearly completed C-I theory’s second draft and decided to include this short section to justify the juxtaposition of several information-related sciences and ethics. It was both gratifying and encouraging to see that another author with a post graduate education in both philosophy and a science had come to some of the same conclusions that I had – an instance of convergence, where the same problem or question provokes similar solutions and answers, e.g., all animals that have true flight have evolved wings. A few of the convergences between this book and Harris’ include:

- The well-being of things, including our brain, is a common basis for that which we positively value.

- Well-being is often secondary to the physical nature of a system, which is, in part, something that science works to better understand.
- There is more than one way for something like a society to flourish, but there are many more ways for it to do poorly.
- That a science of morality will have many of the features of utilitarianism and virtue ethics.

There are more similarities that I will not recount, while substantial differences in conclusions are few and separate from our primary theses. For example, he does not believe in free will, whereas C-I theory is agnostic to silent on this issue. The biggest difference, however, is that *The Moral Landscape* primarily emphasizes how science in general should be able to offer guidance for many of the issues of ethics. *Complex-Information (C-I) theory* is more specific in its goal of describing how several information related sciences can offer valuable insights into key ethical issues, especially its axiology. Therefore, the two publications should be seen as being complimentary rather than largely overlapping.

Of course, Harris' treatise and C-I theory are far from the first serious attempt to combine science and ethics. Jeremy Bentham had proposed that a scientific approach to utilitarianism could help to make happiness more determinative (Bentham, J., 2005). The Polish philosopher Maria Ossowska (1896-1974) believed that sociology could be brought to bear on ethics (Zdrenka, M.T., 2006), and Patricia Churchland (1943 -) wished to do the same with neuroscience (Churchland, P.S., 2011). Finally, but not comprehensively, evolution and anthropology bring their respective tools and discoveries to bear on how the morality of humans and related species not only came to be but were almost inevitable with the passage of time, pressures of natural selection, and cultural context (e.g., Kaufman, W., 2013; Laidlaw, J., 2023). Admittedly, however, sciences like sociology and neuroscience seem to be more proximate to the project of ethics than a fundamental physical science like thermodynamics. As my skeptical colleague believed above, many philosophers will likely think that physics has little to offer philosophy outside of ontological and metaphysical considerations.

Additionally, resistance to the juxtaposition of philosophy and science is puzzling because of an inescapable historical reason. Before the mid-1800's the study of nature was called "natural philosophy" as exemplified by Isaac Newton's (1643-1727), *Philosophiae Naturalis Principia Mathematica*, possibly the most important single book in the history of science. Philosophy and science's primary epistemic methods began to diverge even earlier with the writings Sir Francis Bacon (1561-1626) and Renee Descartes (1596-1650), as well as the experiments of Galileo Galilei (1564 -1642) (Brading, K., Marius, S., 2021). As different disciplines over the course of time developed the ability to test proposals empirically, i.e., with controlled or repeated careful observations, the sciences, as they came to be understood, split off from the more purely rational works of philosophy. For example, religion-based philosophy might propose that the planets must move in perfect circles because God prefers geometric perfection, and a secular philosopher might conjecture that all materials are made from earth, wind, water, and fire. These philosophies and their accompanying explicata would have to be abandoned by the learned community as astronomy's observations and calculations demonstrated elliptical planetary orbits, and physics and chemistry revealed that materials were constituted of elements and their combinations, respectively.

On the other hand, philosophy has anticipated errors of science including serious missteps in the ethical use of experimental subjects (e.g., the Tuskegee Syphilis Trials) or offer insights regarding how to differentiate true science from that which is not, e.g., “pseudoscience” (e.g., Popper, K., 1962).

Of course, reason works inextricably with empiricism in science, and experimental philosophy combines reason and empiricism as well (Knobe, J., Nichols, S., 2017). The difference in epistemic methods of the fields is a matter of degree rather than kind. In the admittedly dry lexicon of information, both the sciences and some humanities like philosophy and history work to better comprehend the various constituents of the universe and how they have interacted, are interacting, and will interact.

Nevertheless, I worry that many readers will skip the first three chapters of this book for fear of science’s differing lexicon, methods, and occasional use of mathematics. However, even if mathematics is a distant unfond memory for some, a qualitative and approximate understanding of the precepts will suffice. Even in the practice of medicine, of which I am most familiar, qualitative approximations are usually “good enough.” For example, the drug doses for patients are often not quantitatively calculated as “milligrams per kilogram,” but more roughly determined by whether the patient is an infant, child or an adult. There are exceptions where more accurate dosing of medications are required, of course, but usually the complexities of physiology and variations in metabolic enzymes make exact drug dosing determinations almost superfluous. Similarly, qualitative approximations will work, and quite often cannot be avoided, for most of the physical and informational topics that we will be discussing.

Finally, I hope that the first three chapters, and indeed, the entirety of the book heeds the admonition of the late physicist, Richard Feynman (1918 – 1988): “If you can’t explain something to a first-year student [in college], then you haven’t really understood it” (Colson, R.S., 2014). Similarly, I attempted to avoid excessive scientific jargon, and if I could not, then explained it in a way that is comprehensible to the educated reader. Admittedly, I have not gone deeply into the scientific methods and mathematics (the language of patterns and relationships) that led to many of the discoveries presented here. The genius of the likes of Ludwig Boltzmann, Willard Gibbs, Claude Shannon, and others is unfortunately underemphasized by not describing their methods and mathematics in any detail. Similarly, I fail in adequately acknowledging the genius in contemplative rationale by philosophers like Aristotle (384 BCE - 322 BCE), Immanuel Kant (1724 – 1804), John Stuart Mill (1806 – 1873), John Rawls (1921 - 2002) and others. In brief, the academic “silos” that all too often separate even closely related disciplines is to be despaired because in the end, we are all trying to better comprehend some fantastic aspect of this ultimately ineffable universe, including what we *ought* to do and what we *ought not* do. Despite this apologia for a multidisciplinary approach to something so broad and important as ethics, there will undoubtedly remain those who believe that the silos of science and those of philosophy are only fragmentarily connected by tenuous bridges. If so, to pursue a metaphor that I will employ in the upcoming chapters, hopefully I will convince the reader that the pylons that help to support a bridge between science and philosophy will give us a novel and nontrivial perspective on some important aspects of ethics.

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Chapter 1

A Thermodynamics Grounding for Ethics

Introduction

For many, if not most interested in philosophy, the juxtaposition of the physics subdiscipline of thermodynamics and metaethics or ethics will likely sound strange. How is it that a subfield of physics, whose Greek translation literally means “heat movement,” has any relevance to ethical and even metaethical considerations? While many questions in ethics, metaethics, and philosophy in general defy even our deepest and most conjectural extensions of physics (e.g., physics cannot explain our own consciousness), it is still a worthwhile endeavor to see what thermodynamics, a subdiscipline of physics, might have to offer. Thermodynamics’ laws govern inescapable aspects of how matter and energy *must* behave. Therefore, gleaning what we can from this science might have significant implications for many of the questions with which ethics grapples. At the very least, philosophical discourse faces perils if it contradicts what science has learned with a high degree of certainty.

Thermodynamics – A Brief Overview

Thermodynamics is one of the longest and most well established subdisciplines of physics. With its roots beginning in earnest in the middle of the 19th century, thermodynamics explains many of the fundamental properties and relationships of heat, work, temperature, and energy of large numbers of particles (Drake, G.W.F., 2025; Schroeder, D.V., 2021). The term "thermodynamics" is Greek for "heat movement," which was the primary topic of concern in its early days when steam engines were the highest level of technology. Despite the marked advancements in technology since then, thermodynamics remains relevant today. Fortunately, a basic qualitative understanding of thermodynamics will suffice for the purposes of building a basis for "Complex-Information theory" (C-I theory) that is described in chapter 4. Hence, this chapter primarily gleans information about thermodynamics from several introductory educational resources. (Atkins, P., 2010; Schroeder, D.V., 2021; Grossman, J.C., 2014; Pauken, M., 2011). We do not need, for example, to delve into thermodynamics' many implications for engineering, its deeper mathematics, controversies regarding peripheral details in the field like the entropy of black holes, etc. Instead, having a coherent, mainstream understanding of "classical" thermodynamics, which is what applies at the level of everyday experience, will suffice for beginning to set a foundation for C-I theory.

There are 4 laws that are at the basis of thermodynamics, and it is the 2nd law which is most important for our purposes. However, having a passing acquaintance of thermodynamics other three laws help us to better understand thermodynamics' overall context, so it is worthwhile for us to look briefly at them before delving into the 2nd law. Before we can even begin to discuss the laws of thermodynamics, we must have a working definition of a few thermodynamic related terms.

A thermodynamic "system" is *a group of materials or particles, or simply a space of interest, with definable boundaries that exhibit thermodynamic properties* (Grossman, J., 2014, p.26). A system for thermodynamic purposes can be as small (or smaller) and as simple as a teaspoon of water, or as large and complex as the universe itself. The affected space around a system is called its "surroundings," although in the case of the of the universe, there would be no surroundings as per our current mainstream understanding of the cosmos (Betz, E., 2023).

A few other thermodynamic definitions are also necessary, including the different types of systems that are differentiated by how they interact with their surroundings (Figure 1.1):

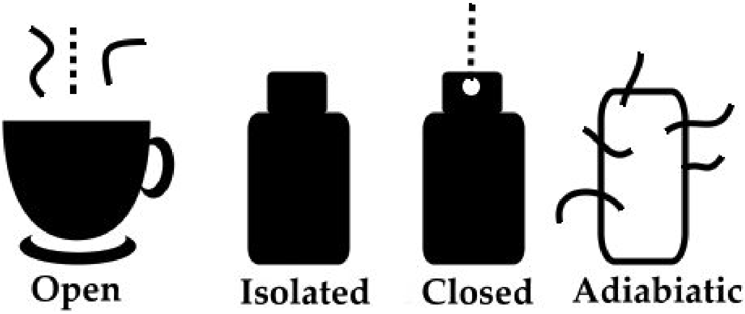

Figure 1.1 *Depictions of the four main types of thermodynamics systems. From left to right: "open," "isolated," "closed," and "adiabatic." A dotted line represents particles leaving or entering a system. Curved lines represent energy leaving or entering a system.*

- As depicted by the open cup in the leftmost illustration, an "open" system allows the exchange of both energy and materials with its surroundings. The wavy lines represent thermal energy being released and the dotted line represents the release of water molecules. Ecosystems and living organisms are biologic examples of open systems.
- An "isolated" system as shown in the second from the left drawing, does not allow the exchange of energy or materi-

als with its surroundings. A sealed thermos bottle approximates an isolated system.

- A “closed” system exchanges material with its surroundings, but not energy. An insulated thermos bottle with a small opening that allows for evaporation of water molecules approximates a closed system.
- An “adiabatic” system allows for the exchange of energy to the surroundings, but not material. A sealed container with thin, conductive walls, like an unopened aluminum can, approximates such an adiabatic system.
- A system in “equilibrium" is equal to its surroundings regarding some parameter. For example, two objects in thermal equilibrium with each other do not have any net change in the kinetic energy of their molecules, i.e., they are at the same temperature.
- A system in “disequilibrium” is not in balance regarding some parameter(s) with its surroundings. Living organisms, for example, must ingest food and often oxygen to maintain various gradients (e.g., material, electrical, thermal) that differ from their surrounding environment. A dead organism can no longer maintain these gradients and subsequently loses its ability to maintain its disequilibrium with its surroundings.

With these necessary definitions in place, we can briefly describe thermodynamics four laws.

The 0th law, whose unusual name is due to historical circumstance, states that, *“if ‘A’ is in thermal equilibrium with ‘B,’ and ‘B’ is in thermal equilibrium with ‘C,’ then ‘C’ will be in thermal equilibrium with ‘A’.”* (Atkins P., 2010, p.5). In mathematics, it is the equivalent of stating that if A=B, and B=C, then A=C, which seems to be a trivial statement. However, thermodynamics is a science that carefully defines and accounts for its terms including "temperature." One immediate implication is that a thermometer (system "B") can be used to determine if two other systems ("A" and "C") are at the same temperature. Temperature in turn is dependent on the kinetic energy (energy from motion) of a system's atoms or molecules (Grossman, J., 2010, p.20), even if that motion is a molecule's vibrations as occurs in a solid. Another important implication of the 0th law is that two systems in contact with each other will not have a net exchange of thermal energy if they are at the same temperature.

The 1st law of thermodynamics, which was actually formulated before the 0th law, is the "conservation of energy." In brief, it states that energy can be neither created nor destroyed in any isolated system, it can only change form, e.g., chemical energy to electrical energy as occurs in batteries or in our nerve cells. The universe is the ultimate isolated system, so, the amount of total energy present at the Big Bang, will be the same today and as it will be the future – even though the type and "quality" (see the second law below) of energy can change over time. What energy *is* remains an ontological mystery to science, but it has a well-known functional definition: *"energy" = the capacity to do work*, and "work" is defined as the *motion against an opposing force* (Atkins, P., 2010, p.16). One of the 1st law's pragmatic implications is that perpetual motion machines are impossible because some of the work energy used by the machine will inevitably be lost to its surroundings as wasted heat. Hence, the loss of heat energy to the machine's surroundings means that that amount of energy is no longer available to allow the machine to continue to work without energy being added to it from without.

The 3rd law of thermodynamics is the furthest removed from everyday experience: *"the entropy of a perfect* (no defects) *crystal is zero when the temperature of the crystal is equal to absolute zero"* (Pauken, M., 2011, p.37). Absolute zero is not theoretically achievable, but scientists have gotten within a few trillionths of a degree from it (Luntz, S., 2021). At absolute 0, all particles are at their "ground state" (lowest) energy level and exhibit no kinetic energy, i.e., movement.

Because of the wide range of contexts in which thermodynamic laws apply, there are other articulations of the 0th, 1st, and 3rd laws. Unfortunately, the diversity of definitions especially applies to the 2nd law, which is most important for the foundation of C-I theory.

The Second Law of Thermodynamics and Entropy

The second law of thermodynamics is considered by many to be one of the most, if not *the* most, important concepts in science. Peter Atkins, a physicist at Oxford University, equates understanding entropy in science to being familiar with the works of Shakespeare in the humanities (Atkins, P., 2010). Sir Arthur Eddington (1882-1944), an early 20th century physicist and philosopher of science, also opined the following regarding the resolute soundness of thermodynamics 2nd law (Malizia, D., Tarsitani, C., 1995, p. 355):

> The law that "entropy" always increases—the Second Law of Thermodynamics—holds, I think, the supreme position among the laws of Nature. If someone points out to you that your pet theory of the universe is in disagreement with Maxwell's equations—then so much the worse for Maxwell's equations. If it is found to be contradicted by observation—well these experimentalists do bungle things

> sometimes. But if your theory is found to be against the second law of thermodynamics, I can give you no hope; there is nothing for it but to collapse in deepest humiliation.

We can more fully restate Eddington's brief description of the second law above as, "entropy always increases in the universe over the course of time." Even if Eddington's description of the 2^{nd} law above can be refined the slightest bit, we must understand what "entropy" means before we can proceed. Entropy is most often described as being "the degree of disorder of a system." This description approximates what it means and is sufficient for most mundane purposes. However, some physicists point out that crushed ice in a container appears less ordered than seemingly homogenous liquid water in container. Therefore, sometimes order is a human perception that does not necessarily reflect the deeper or microscopic state of a system (Schroeder, D., 2021, p.75). Physicists also use other terms to describe entropy, including a system's measure of "dispersion," "smoothness," "homogeneity," "multiplicity" and "degrees of freedom" (Grossman, J.,2010, p.30; Schroeder, D., 2021). For our purposes, a system's degree of *randomness* of its thermodynamic parameters will suffice because it seems to strike the right balance between rigorousness and descriptiveness. (Admittedly, I often resort to using the simpler term "disorder" in this book as well.) Figure 1.2 illustrates in part what a physicist means by the randomness of energy levels of a system's particles. At low entropy, the particles jump in and out of 3 different energy states. At higher entropy, the system's various particles (which are vast in number) jump in and out of more different energy levels (Atkins, P., 2010). For simplicity's sake, only 3 versus 5 energy levels are depicted in the diagram, although there are typically many more. Additional randomness in energy can be attained by different types of atomic or molecular movements as shown in figure 1.3. For example, the energy of particles in a solid will vibrate more vigorously at their positions within the materials lattice. Particles in a liquid can rotate as well as vibrate more vigorously, and gases can move through space or "translate" in addition to the other movements. Hence, as a material changes it "phases" from solid to liquid to gas, the randomness in *types* of movement (vibrational, rotational, translational) increase as well.

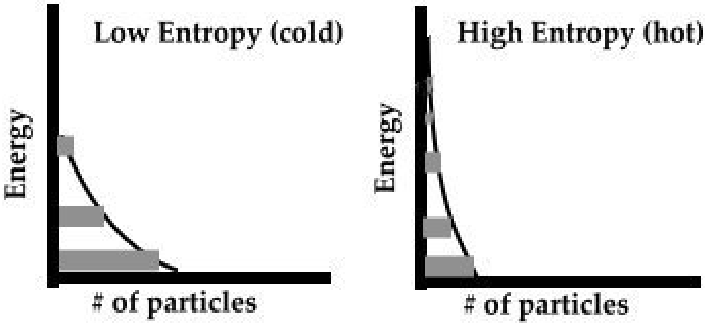

Figure 1.2 *Particles in a colder or lower entropic system will occupy fewer energy levels than a hotter or higher entropic system.*

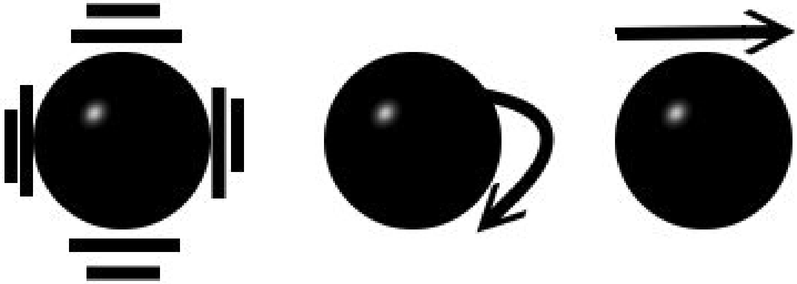

Figure 1.3 *From left to right, particles with vibrational, rotational, and translational motion. Particles of solids only have vibrational movement, particles in liquids can have vibrational and rotational, while particles in a gas have all three.*

The particles can also gain an increase in randomness (greater entropy) by becoming more dispersed in space as when you spray a fragrance from a bottle into an open room or allow an open cup of coffee to evaporate as in the leftmost illustration in figure 1.1. What we should gain from figures 1.1 to 1.3 is a mental picture of what a key thermodynamic term like “entropy” means at the microscopic level, if only in a qualitative way.

Thermodynamic-centric descriptions of entropy are important for engineers, chemists, physicists, and other fields because of its implications regarding matter, energy and their uses. Another common description of entropy is that it is the amount of energy in a system that is not available for work, where “work” is the defined as the potential ability to move something. Entropy is the energy that will be “wasted” as heat. “Free” energy is that portion that performs the work. Another common explanation regarding entropy is that whereas the 1st law of thermodynamics states that the *quantity* of energy in a system does not change, the 2nd law describes the *quality* of a system’s energy content. A system with a high degree of energy quality will have a low degree of entropy and be capable of performing more work. A system with a low degree of energy quality will have high entropy and be capable of performing less work, even if the quantity of energy is the same. The main source of high-quality energy for the Earth’s processes, is the Sun. The Earth intercepts a tiny fraction of the Sun’s high frequency photons (high quality, lower entropy energy), which then provides energy for everything from photosynthesis to weather. Earth then emits lower frequency photons back into space (low quality, higher entropy energy) (Chaisson, E.J., 2001).

To contrast a system that releases high entropy or low-quality energy versus low entropy or high-quality energy, imagine an open pot of boiling water versus the same pot that has a lid with a small hole in it. If the pot has no lid, the resulting steam will be released as highly randomized thermal energy with no work done except for the small amount done to push the atmosphere's molecules aside. The pot of boiling water with a hole in the lid has a directed, less randomized release of energy and can move something besides atmospheric molecules– say a piston – much more effectively. This latter example is the essence of a basic steam engine that was first invented by Thomas Savery in 1698 (Editors of Encyclopedia Britannica, 2024). As others have pointed out, the use of fire to create randomized thermal energy for keeping us warm and cooking food might have been invented as early as 1-1.5 million years ago by *Homo erectus*, a predecessor to *H. sapiens* (Gowlett, J.A.J., 2016). Learning how to harness substantial free energy or work from fire, however, began only a little over 300 years ago with the invention of the steam engine.

A corollary of the limitations of a steam engine, and indeed any machine designed to perform work, is that at least some of the energy will be lost to entropy or thermal energy (a.k.a. ~heat), to its surrounding – not all energy can be converted to the work. That is, consistent with the 2nd law, any real-world process must increase entropy, and consequently no engine can be 100% efficient. Some of that jet of steam will heat the surroundings rather than pushing on the piston. This limitation applies to any other real-world engine no matter how cleverly contrived.

The thermodynamic implications of entropy affect analyses in other areas besides steam engine engineering. For example, a chemist might want to know if the entropy increases in a chemical reaction because if so, the reaction will result in a net surplus of energy. Biologists wonder how life, systems with relatively low entropy, came into existence despite the great majority of the rest of the universe marching towards higher states of entropy. Considerations of entropy can even be as arcane as to make astrophysicists ponder whether cosmic black holes exhibit entropy. If so, entropy, or the dispersion of particles into the surroundings will cause them to essentially evaporate over vast periods of time as was proposed by the late physicist, Stephen Hawking (1942-2018) (Mann, A., 2023).

For C-I theory, however, we will find it most useful to switch to a "statistical mechanical" or information-centric definition of entropy: *"Entropy" is the number of microstates of a system's macrostate* (Grossman J.C., 2014 p48). At this point it might seem that the last thing we need is yet another definition of "entropy." However, the other definitions of entropy are descriptive or explain some of its energetic consequences, rather than being definitional. The brief definition above warrants the explanation of its key terms and a brief recounting of its origin.

In the later 1800's the Austrian physicist, Ludwig Boltzmann (1844-1906), studied macroscopic (i.e., measurable or observable by the naked eye) thermodynamic phenomena like the equilibration of temperature of two adiabatic systems in contact with each other (i.e., they exchange energy, but not matter). He determined that their subsequent thermal equilibration could be explained by assuming that it was the *statistical averaging* of the kinetic energy of atomic particles in each system that explained the process. In other words, many observable or measurable "macrostate" properties of a system, like temperature and pressure, were determined by the average behavior of its unobservable microscopic constituents at that moment – its "microstate." Importantly, you did not need to try to account for the behavior (e.g., motion, location) of each of the untold quadrillions of individual particles that constitute even a miniscule observable system. Instead, you could use mathematical statistics to understand their collective behavior (Goldstein, S., 2001). Boltzmann's key insight is all the more remarkable because the existence of atoms had not yet been agreed upon by most of the broader scientific community. Some science historians conjecture that the resistance to his insights by his peers contributed to his suicide in 1906 (Bausek, N., Washietl, S., 2018) – just a year after the publication of Einstein's paper of how Brownian motion (e.g., the wiggling of dust particles on a surface of water) strongly supported the existence of moving atoms. Unfortunately for Boltzmann, it took a few years for Einstein's discovery to be disseminated, replicated, and accepted by the scientific community. After all, at the time of the publication of his paper on Brownian motion in 1905, Einstein was "just a patent clerk."

Ironically, Boltzmann's tombstone is famously inscribed with the equation that describes how to determine the number of microstates of a system's macrostate: $S = k \log W$, where "S" is entropy, "k" is the Boltzmann constant in units of energy per degree Kelvin, and "W" is the number of microstates of a system expressed as a logarithm. In most instances, we can avoid mathematics and remain with a qualitative description and rough approximations for grounding C-I theory. This equation is important, however, because it helps to substantiate the close relationship between entropy and syntactical information, as we will see in the next chapter.

For this equation and others that include entropy, "S" was chosen by early physicists to stand for a system's entropy, simply because most of the other letters had already been assigned to represent other thermodynamic parameters, e.g., "E" stands for "energy." Because "k" is a fixed number and we are not concerned with precise quantitative values, we can essentially ignore it for our purposes. “W,” the number of a system’s microstates, is preceded by the “log” or logarithm function. If you do not recall from your earlier years in mathematics or perhaps never learned it, logarithms simply make it much easier to manage huge numbers, like those of a typical system’s possible microscopic number of arrangements or microstates. If a system, for example, has 1,000,000,000,000,000,000 (10^{18}) microstates, then its “base 10” logarithm (shown formally as $\log_{10}$) is the “power” to which 10 was raised to arrive at the number, or in this case, 18. Logarithms also make it easier to add and multiply huge numbers. Note that you will also often find the formula “$S = k \ln W$” or “$S = k \ln \Omega$” in some references. “ln” stands for “natural logarithm” and has the constant “Euler’s number” (~2.71828) as its base ($\log_{2.71828}$) and “Ω” is just a substitute for “W.” Fortunately, for our purposes the base used for the logarithm is not important and does not affect the concepts with which we are concerned. (If you are curious to learn a bit more or refresh your basic logarithm functions, see the Appendix.) In brief, the possible number of “microstates” of a system also reflects the degree of randomness of its particles that was described earlier.

The formula $S = k \log W$ works for systems where the probability of each microstate is equally likely to occur. For instance, in a glass of pure water, it is as likely for H2O molecules in any given volume to be in microstate “A” as it is in microstate “B.” Willard Gibbs (1839-1903), another giant in thermodynamics, however, determined a more generalized equation where the different components of a system have a different probability of being in a particular microstate; perhaps because one liquid substance in a mixture has a greater quantity than another, so that it is more likely to be present in any particular microstate of the liquid. For this kind of situation, entropy is determined by the equation, $S = k \Sigma_1^m p_i \log 1/p_i$, where Σ_1^m is the summation of the number of microstates, "m," with values of p_i. "p_i" is the probability of a component in the system being in a particular microstate (Stone, J.V.,2015, p.175). Importantly, this more comprehensive equation reduces to $S = k \log W$ if all the microstates have the same probability of occurring. Again, the mathematics is not as important for our purposes but presented here to later demonstrate the close relationship between entropy and information.

Before we move to the next metaphorical pylon (information) of the bridge spanning the physical world to C-I theory. It will be helpful to introduce another key term – "phase space." Phase space is an abstract representation of the number of microstates (~ degree of randomness) that comprise a system in a particular macrostate. Therefore, the larger the phase space, the larger the number of microstates and the greater the entropy of a system. Each point within a phase space abstractly represents a particular microstate that is consistent with that macrostate (e.g., a system with a particular volume, temperature, pressure, etc.). As a simple example, figure 1.4 shows a hypothetical phase space for an iron rod (a simple thermodynamic system) of a given size. We can increase the phase space of the rod's entropy in innumerable ways as shown in part by the bars on the right. The top bar is heated, and its atoms are vibrating at more different energy levels. The middle bar is broken into pieces which disperse its particles. The bottom bar is beginning to oxidize (rust), which releases heat because it is an exothermic reaction, and increases the variety of its compositional particles. These examples illustrate just three ways that processes can increase a system's entropy, number of microstates, and the coinciding area of its phase space. Note that if more accurately depicted, phase space 2 would many magnitudes larger than shown.

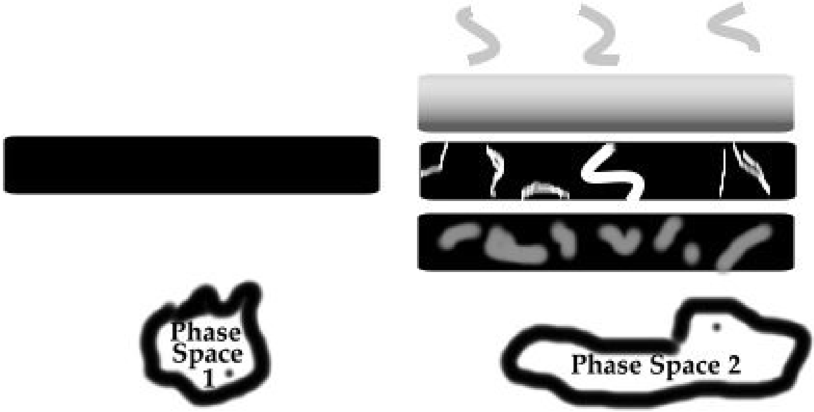

Figure 1.4 *The left rectangle represents an iron bar at room temperature and the phase space is an abstract area where each point in it represents a possible microstate or arrangement & properties of its particles. The rectangle at top right shows a heated iron rod above room temperature, the middle bar one that is broken into pieces, and the bottom bar is rusting. The phase spaces of the right bars are larger than the phase space of the left bar because there are more possible microstates. If shown to actual scale, "phase space 2" would be many magnitudes larger than "phase space 1."*

Why Does Entropy Always (Essentially) Increase?

In a few circumstances, entropy can remain unchanged as it does in an isolated system (no exchange of matter or energy with its surroundings) or if the system is in a full state of equilibrium with its surroundings. Entropy can also remain unchanged with "reversible" processes. where "infinitesimal" changes occur. An example of an infinitesimal change is a system that slowly undergoes a miniscule alteration in one of its parameters like pressure or temperature and then reverses to its original state (Schroeder, D.V., 2021, p.76). Reversible processes, however, are ideal and theoretical. In the "real" world, processes are irreversible and do not revert to their prior state without work energy being expended and entropy increased somewhere else in the surroundings. In other words, in the real world, Eddington's claim that the universe only increases in entropy is safe as he stated it would be.

The one-way nature and inevitably of the universe's increase in entropy has led physicists to proclaim that entropy is "the arrow of time"- both time and entropy only proceed in one direction. Entropy's ubiquity also begs the question whether there is an entropic "force" like gravity or electromagnetism? However, no force of entropy has ever been discovered (Müller, I., 2007; Roos, N., 2014). Instead, entropy is "driven" by straightforward and overwhelming statistics - there are *many, many* more randomized microstates for macroscopic systems than the contrary. Hence, iron bars next to each other equalize their temperatures if there is a difference, gas quickly disseminates throughout a room, cars will fall apart and rust over time, and water evaporates as vapor. (Scientists believe that over vast periods of time, even an isolated system will eventually decay to become a scattering of photons.) Although it is not *impossible* for one iron bar to spontaneously become warmer than its surroundings, air to concentrate into part of a room, a car to become less rusted, or liquid water to reform from vapor, the lower entropic states are so statistically unlikely that active stars will cease to exist in about a quintillion (10^{18}) years or more before any of these events have a reasonable chance of occurring (Adler, D., 2023) .

The analogies used by various authors to illustrate the statistics relevant to entropy are legion - from an untended teen's bedroom becoming messy over time, to cream mixing with coffee, to the possible arrangements of black and white stones on the East Asian board game of "go." However, to begin to have a basic grasp of the statistics relevant to entropy, I will begin by using a scenario similar to the flipping of 3 different coins as explained by the physicist Daniel V. Schroeder (Schroeder D.V., 2021, pp 49-51). Table 1.1 uses a more illustrative concern regarding whether someone could suffocate because all the oxygen ended up accidentally on the other side of a room.

Left side of room	Right side of room
O1 O2 O3	
O1 O2	O3
O1	O2 O3
O1 O3	O2
O2 O3	O1
O3	O2 O3
O2	O1 O3
	O1 O2 O3

Table 1.1 *Each "O" represents an oxygen molecule. Because atoms and molecules are indistinguishable from each other, breathing in O1 is the same as O2 or O3. There are only 2 out of 8, or 1 out of 4 (25%) situations or "macrostates" where there is no oxygen in half of the room. Conversely, there are 6 out 8, or 3 out 4 (75%) macrostates where there is oxygen in each side of the room. With the huge numbers of oxygen molecules in a real room, the chance of being in a "no" or "low" oxygen macrostate area of the room dramatically diminishes so that the chance of suffocating approaches zero even over the course of trillions of years.*

For simplicity's sake, table 1.1 shows only 3 oxygen molecules that could end up in one side of the room or the other. The chance of each microstate is 1 in 8 or 12.5% depending on how you prefer to express your chances. Therefore, there is a 12.5% chance of suffocation because all the oxygen is gone from your side. The chance of all the oxygen being in either side of the room, however, is 1 in 4 or 25%. Additionally, there are a total of 6 ways to have 2 molecules on one side and 1 molecule on the other side of the room. Therefore, statistically it is three times more likely for the molecules to be more evenly spread through the room.

Of course, in this example the odds of suffocation are not trivial. There is a 1 in 8 chance of suddenly suffocating. However, real-life macroscopic systems have *vast* numbers of molecules and the likelihood of the molecules being almost evenly distributed throughout the room increases many, many fold, while the likelihood of unusual arrangements, like all or even most being to one side of the room, decreases many, many fold and approaches zero.

To begin to illustrate how larger numbers of particles immensely diminish the odds of entropy spontaneously decreasing, consider the results of the relevant mathematics. The number of possible arrangements of the constituents a system can be in, is determined by a mathematical function called the "factorial." A factorial function is indicated by the exclamation mark "!" and is calculated by multiplying each number by the next smaller integer down to 1. Therefore, 5! = 5x4x3x2x1 = 120. The appropriateness of using the "!" symbol becomes apparent when you discover that seemingly small factorial calculations result in huge numbers. For example, 60! calculates to about 8.32×10^{81} possibilities. This quantity is about the same as the number of atoms in the observable universe (Baker H, 2021). Entropy is about statistical probability, and the probability of 60 molecules all being on one side of a space is not far from picking the "right" two atoms from the vastness of the observable universe. Furthermore, because a small 3m x3m x 3m room has about 2.7×10^{25} air molecules (Benson, T., 2021), the likelihood of them not being evenly distributed around the room is vastly more unlikely than the only 60 molecules mentioned above. Thus, the claim that the stars will burn out before it is likely that all