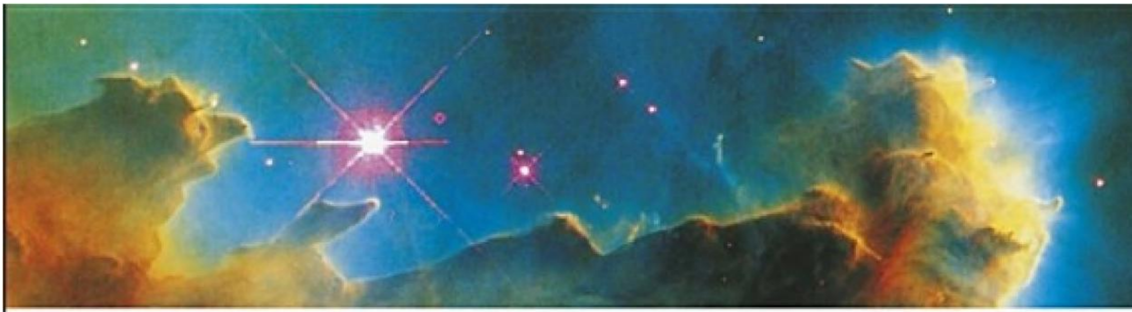
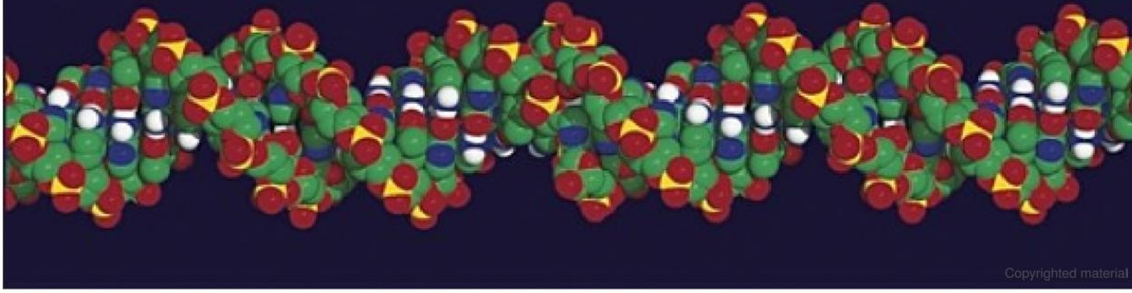


ALMOST EVERYONE'S GUIDE TO SCIENCE

JOHN GRIBBIN



CONTENTS

SCIENTIFIC NOTATION

When dealing with very large or small numbers, as we shall be in this book, it is convenient to use scientific shorthand to avoid writing out strings of zeroes. In this standard scientific notation, 10^2 means 100 (a 1 followed by two zeroes), 10^3 means 1000, and so on. 10^{-1} means 0.1, 10^{-2} means 0.01, and so on. This notation comes into its own when we are dealing with numbers like Avogadro's Number (see chapter one), which is written as 6×10^{23} in scientific notation, shorthand for 600,000,000,000,000,000,000,000.

One point to watch is the significance of making a seemingly small change in the power of ten involved. 10^{24}, for example, is ten times bigger than 10^{23}, and 10^6 is not half of 10^{12}, but one millionth (10^{-6} times) as big.

We also follow the scientific convention in which a billion is a thousand million, or 10^9.

INTRODUCTION

IF IT DISAGREES WITH EXPERIMENT IT IS WRONG

The fate of specialists in any one area of science is to focus more and more narrowly on their special topic, learning more and more about less and less, until eventually they end up knowing everything about nothing.

It was in order to avoid such a fate that, many years ago, I chose to become a writer about science, rather than a scientific researcher. The opportunity this gave me to question real scientists about their work, and to report my findings in a series of books and articles, enabled me to learn less and less about more and more, although as yet I have not quite reached the stage of knowing nothing about everything. After thirty years of this, and many books focusing on specific aspects of science, it seemed a good idea to write a general book, giving a broad overview of science, while I am still at the stage of knowing a little bit about most things scientific.

Usually, when I write a book the target audience is myself – I write the book about, say, quantum physics, or evolution, that I wish somebody else had written for me so that I would not have had to go to the trouble of finding things out for myself. This time I am writing for everybody else, in the hope that there will be something here for almost everyone to enjoy. If you know a little quantum physics (or even a lot), you may find here something you didn't know about evolution; if you know about evolution, you may find something new to you about the Big Bang, and so on.

So, although I am aware of the ghost of Isaac Asimov looking over my shoulder (I hope with approval) at such a wide-ranging project, this is not 'John Gribbin's Guide to Science', but a guide for almost everyone else. A guide not so much for fans of science and the cognoscenti but more a guide for the perplexed – anyone who is vaguely aware that science is important, and might even be interesting, but is usually scared off by the technical detail. You won’t find such technicalities here (they have all been removed by my co-author, who has kept my wilder scientific extravagances in check and has ensured that what remains is intelligible to a layperson). What you will find is one person’s view of how science stands at the end of the twentieth century, and how the different pieces fit together to produce a coherent, broad picture of the Universe and everything it contains.

The fact that the pieces do fit together in this way is something you might miss from focusing too closely on one aspect of science, such as the Big Bang or evolution, but it is an extremely important feature of science. Both evolution and the Big Bang (and all the rest) are based on the same principles, and you can’t pick and choose which bits of the scientific story you are going to accept.

I often receive communications from people who, for one reason or another, cannot accept the special theory of relativity, which tells us that moving clocks run slow and moving rulers shrink. Sometimes these people struggle desperately to find a way around this, while still accepting everything else in science. But you cannot. The special theory of relativity does not stand in isolation, as a theory about moving clocks and rulers, but comes into our understanding of, for example, the way mass is converted into energy to keep the Sun shining, and how electrons behave inside atoms. If you threw away the bits of the theory that seem to refute common sense, you would be left with no explanation of why the Sun is shining or of the periodic table of the elements. And this is just one example.

As I hope this book will make clear, *everything fits together* in the modern scientific world view. This scientific world view is the greatest achievement of the human intellect, and the power of that achievement stands out more clearly from a look at the broad picture than it does from too close attention to any one detail.

There are two remarkable, interconnected features of the scientific world view that are often overlooked, but are well worth pointing out. The whole thing has taken only about four hundred years to develop (starting from the time of Galileo, which seems as good a moment as any from which to date the beginning of modern scientific enquiry). And it can all be understood by a single human mind. Maybe we cannot all understand every bit of the scientific world view; but quite a few individual human beings can, even though people have such limited lifespans. And although it may take a genius to come up with an idea like the theory of evolution by natural selection, once that idea is formulated it can be explained to people of average intelligence – often provoking the initial response, ‘How obvious; how stupid of me not to have worked that out for myself.’ (This was, for example, pretty much Thomas Henry Huxley’s reaction when he first read Charles Darwin’s *Origin of Species*.) As Albert Einstein said in 1936, ‘The eternal mystery of the world is its comprehensibility.’

The reason why the Universe is comprehensible to mortal minds is that it is governed by a small set of very simple rules. Ernest Rutherford, the physicist who gave us the nuclear model of the atom early in the twentieth century, once said that ‘science is divided into two categories, physics and stamp-collecting’. He was not being entirely facetious, although he did have a genuine disdain for the other sciences, which made the fact that he was awarded the Nobel Prize for Chemistry (in 1908, for his work on radioactivity) deliciously inappropriate. Physics is the most fundamental of the sciences, both because it deals most directly with the simple rules that govern the Universe and the simple particles that everything in the Universe is made of, and because the methods of physics provide the archetype used by other sciences in developing their own parts of the world picture.

The most important of these methods is the use of what physicists call models. But even some physicists do not always appreciate just what their use of models is really all about so it is worth spelling this out before we think about applying the technique.

To a physicist, a model is a combination of a mental image of what some fundamental (or not so fundamental) entity is like, and a set of mathematical equations that describe its behaviour. For example, one model of the air that fills the room in which I am writing these words would regard every molecule of gas in the air as like a tiny, hard ball. There are accompanying equations which describe, at one level, how those little balls collide with one another and bounce off each other and the walls of the room and, at another level, how the average behaviour of very many of those little hard balls produces the pressure of the air in my room.

Don't worry about the equations – I shall largely ignore them in this book. But remember that good models always include the equations, and it is the equations that people work with in order to make predictions about the way objects behave – to calculate, perhaps, the way in which the pressure of the air in my room would change, other things being equal, if the temperature went up by ten degrees Celsius. The way to tell a good model from a bad one is to test it by experiment – in this case, warm the room up by ten degrees and see if the new pressure you measure matches the calculated pressure predicted by the model. If it does not, the model at best needs improvement and at worst may have to be discarded altogether.

Richard Feynman, the greatest physicist of the twentieth century, summed up the scientific process like this in a lecture he gave in 1964, using the word 'law' but forcefully making a point which applies equally well to models:

> In general we look for a new law by the following process. First we guess it. Then we compute the consequences of the guess to see what would be implied if this law that we guessed is right. Then we compare the result of the computation to nature, with experiment or experience, compare it directly with observation, to see if it works. If it disagrees with experiment it is wrong. In that simple statement is the key to science. It does not make any difference how beautiful your guess is. It does not make any difference how smart you are, who made the guess, or what his name is – if it disagrees with experiment it is wrong.

That is what science, and scientific models, are all about. *If it disagrees with experiment it is wrong.* But there is a more subtle point. Even if it *does* agree with experiment, that does not mean that a model is 'right' in the sense of being some eternal, universal Deep Truth about the nature of the thing that is being studied. Just because molecules can be treated as little, hard balls for the purpose of calculating the pressure of gas in a room, this does not mean that the molecules *are* little, hard balls – it means that under certain circumstances they behave *as if* they were little, hard balls. Models work within clear-cut – usually – limits, and outside those limits they may have to be replaced by other models.

To make this clear, let's take a different perspective on the image of the molecules of gas in the air in my room. Some of those molecules will be water vapour, and molecules of water, as every schoolchild knows, are made up of three atoms – two hydrogen and one oxygen, written as H_2O. For some purposes, a convenient model of the water molecule is two smallish hard balls (the hydrogen atoms) joined to a single larger hard ball (the oxygen atom) to make a V shape, with the oxygen at the vertex of the V.

For these purposes, the links between the atoms can be regarded as like little stiff springs, so that the atoms in the molecule can jiggle about, vibrating to and fro. This kind of vibration is associated with a characteristic wavelength of radiation – because the atoms carry electric charge (more of this later), if they are forced to vibrate in this way they will radiate microwave radio emission, and, alternatively, if the right kind of microwave radio emission is directed at the molecules they will vibrate in sympathy.

This is exactly what happens in a microwave oven. Microwaves tuned to the wavelengths that make water molecules vibrate fill the oven and make water molecules in the food vibrate, so that they absorb energy and heat up the food. This behaviour is not only seen in the kitchen or the laboratory – it is by studying the microwave radio emission coming from clouds of gas in space that astronomers detect the presence of water molecules in space, along with many other molecules.

So if you are a radio astronomer looking for molecules in space, or an electrical engineer designing a microwave oven, the stick and ball model of a molecule of water is a good one, provided the sticks joining the atoms are a little bit springy. You no longer regard the whole molecule as a single hard sphere; but you do regard the individual atoms, such as the oxygen atoms, as individual hard spheres.

A chemist analysing the composition of a substance would have yet another perspective. If you want to know which kinds of atoms are present in a substance, one way to find out is to study the light that they radiate when they get hot. Different kinds of atoms radiate different colours, very sharply defined lines in the rainbow spectrum of light – one of the most familiar examples is the bright orange-yellow colour of streetlights which contain compounds of sodium. It is the atoms of sodium (in this case, excited by an electric current, rather than by heat) which radiate this particular colour of light.

The model used to describe how this light is produced sees an atom not as a single hard sphere but as a tiny central nucleus (which can itself be thought of as a single hard sphere for now) surrounded by a cloud of tiny, electrically charged particles, called electrons. The central nucleus has a positive electric charge and the electrons each have negative electric charge, so that overall the atom has zero electric charge. The bright lines in the spectrum associated with a particular kind of atom are then explained in terms of the way electrons move in the outer part of the atom. What distinguishes one kind of atom from another, chemically speaking, is the number of electrons (8 for oxygen, just 1 for hydrogen, 11 for sodium); and because each kind of atom has its own unique arrangement of electrons, each kind of atom produces its own unique pattern of coloured lines in the spectrum.

I could go on, but the point is clear already. The model which treats molecules of air as little hard balls is a good one, because it works when you use it to calculate how pressure changes when temperature changes. The model which treats molecules as being made up of smaller hard spheres (atoms) held together like bunches of grapes is also a good one, because it works when you use it to calculate the way vibrating molecules make radio waves. And the model which treats atoms not as indivisible hard spheres but as tiny nuclei, surrounded by clouds of electrons, is also a good one, because it works when you use it to calculate the colour of light associated with a particular kind of atom.

None of the models is the ultimate Deep Truth, but they all have their part to play. They are tools which we use to help our imaginations to get a picture of what is going on, and to calculate things which we can test directly by measurement, such as the pressure of air in a room or the colour of light radiated by a hot substance.

Just as a carpenter would not use a chisel to do the same job as a mallet, so a scientist must choose the right model for the job in hand. When Feynman says ‘if it disagrees with experiment it is wrong’ he means if it disagrees with an *appropriate* experiment. The model of a molecule of water vapour as a single hard sphere does not allow for the possibility of the kind of vibration associated with microwaves, so it ‘predicts’ that water vapour will not radiate microwaves. This means that it is the wrong model to use if we are interested in microwaves. But it does *not* mean that it is the wrong model to use if we are interested in how the pressure of air in the room is affected by increasing the temperature.

Everything in science is about models and predictions, about finding ways to get a picture in your head of how the Universe works, and ways to make calculations that forecast what will happen in certain circumstances. The further we get from the ordinary world of everyday life, whether towards the very small scale or towards the very large scale, the more we have to rely on analogies: an atom is, under certain circumstances, ‘like’ a billiard ball; a black hole is, in some sense, ‘like’ a dent in a trampoline.

It would be tedious to keep qualifying the use of the various models in this way, and now that I have got this off my chest I shall largely refrain from doing so and trust you to remember the qualification that even the best model is only a good one in its own context, and that chisels should never be used to do the job of mallets. Whenever we describe something as being ‘real’, what we mean is that it is the best model to use in the relevant circumstances.

With that proviso, starting out from the scale of atoms, I shall take you down into the world of the very, very small, and then out into the Universe at large, giving the best modern understanding (the best model) of the nature of things on each scale. They are all true, in that they agree with experiment; they all fit together, like pieces in a jigsaw puzzle, to give a coherent picture of how the Universe, and everything in it, works; and it can all be understood, at least in outline, by an average human mind.

There is another feature of science, a view which I hold strongly, and which has shaped the structure of this book (and my entire career), but which is not necessarily shared by all scientists. To me, science is primarily an investigation of our place in the Universe – the place that people occupy in a world which ranges from the tiniest subatomic particles to the furthest reaches of space and time. We do not exist in isolation, and science is a human cultural activity, not a purely dispassionate striving after truth, no matter how hard we might try. It is all about where we came from, and where we are going. And it is the most exciting story ever told.

John Gribbin
December 1997

CHAPTER ONE

ATOMS AND ELEMENTS

In 1962, in a series of lectures given for undergraduates at Caltech, Richard Feynman placed the atomic model at the centre of the scientific understanding of the world. As he expressed it:

> If, in some cataclysm, all of scientific knowledge were to be destroyed, and only one sentence passed on to the next generations of creatures, what statement would contain the most information in the fewest words? I believe it is the *atomic hypothesis* (or the atomic *fact,* or whatever you wish to call it) that *all things are made of atoms – little particles that move around in perpetual motion, attracting each other when they are a little distance apart*, *but repelling upon being squeezed into one another.* In that one sentence, you will see, there is an *enormous* amount of information about the world, if just a little imagination and thinking are applied.

The emphasis is Feynman's, and you can find the whole lecture in his book *Six Easy Pieces.*[1] In the spirit of Feynman, we start our guide to science with atoms. It is often pointed out that the idea of atoms, as the ultimate, indivisible pieces of which matter is composed, goes back to the time of the Ancient Greeks, when, in the fifth century BC, Leucippus of Miletus and his pupil Democritus of Abdera argued the case for such fundamental entities. In fact, although Democritus did give atoms their name (which means 'indivisible'), this is something of a red herring. The idea wasn't taken seriously by their contemporaries, nor by anyone else for more than two thousand years. The real development of the atomic model dates from the end of the eighteenth century, when chemists began the modern investigation of the properties of the elements.

The concept of elements – fundamental substances from which all the complexity of the everyday world is made up – also goes back to the early Greek philosophers, who came up with the idea that everything is made up of different mixtures of four elements – air, earth, fire and water. Apart from the name 'element', and the idea that an element cannot be broken down into any simpler chemical form, there is nothing left of the Greek idea of elements in modern chemistry, which builds from the work of Robert Boyle in the middle of the seventeenth century.

Boyle was the first person to spell out the definition of an element as a substance that could combine with other elements to form compounds, but which could not be broken down into any simpler substance itself. Water, for example, is a compound which can be broken down chemically into its component parts, oxygen and hydrogen. But oxygen and hydrogen are elements, because they cannot be broken down further by chemical means. They are not made of other elements. The number of known elements increased as chemists devised new techniques for breaking compounds up; but by the nineteenth century it was becoming clear which substances really were indivisible.

The breakthrough in understanding the way elements combine to make compounds came when John Dalton revived the idea of atoms at the beginning of the nineteenth century. He based his model on the discovery that for any particular compound, no matter how the compound has been prepared, the ratio of the weights of the different elements present is always the same. For example, in water the ratio of oxygen to hydrogen is always 8:1 by weight; in calcium carbonate (common chalk) the ratio of calcium to carbon to oxygen is always 10:3:12 by weight.

Dalton’s explanation was that each kind of element is composed of one kind of identical atoms, and it is the nature of these atoms that determines the properties of the element. On this picture, the key distinguishing feature that makes it possible to tell one kind of atom from another is its weight. When two or more elements combine, it is actually the atoms of the different elements that join together, to make what are now known as molecules. Each molecule of a compound contains the same number of atoms as every other molecule of the compound, each with the same numbers of atoms of each of the elements involved in each molecule. A molecule of water is made up of two hydrogen atoms and one oxygen atom (H_2O); a molecule of calcium carbonate is made up of one atom of calcium, one atom of carbon, and three atoms of oxygen ($CaCO_3$). And we now know that in some elements the atoms can join together to make molecules, without other elements being involved. The oxygen in the air that we breathe, for example, is made up of di-atomic molecules, O_2 – these are not regarded as compounds.

Dalton's atomic model was a huge success in chemistry, but throughout the nineteenth century some scientists regarded it only as a useful trick, a way to calculate the way elements behaved in chemical reactions, but not a proof that atoms are 'real'. At the same time, other scientists were finding increasingly compelling evidence that atoms could be regarded as real entities, little hard balls that attract each other when some distance apart but repel when pushed together.

One line of attack stemmed from the work of Amadeo Avogadro (who was, incidentally, the person who showed that the combination of atoms in a molecule of water is H_2O, not HO). In 1811 Avogadro published a paper in which he suggested that equal volumes of gas at the same temperature and pressure contain equal numbers of atoms. This was before the idea of molecules was developed, and we would now say that equal volumes of gas at the same temperature and pressure contain equal numbers of molecules. Either way, though, what matters is that Avogadro's model envisaged equal numbers of little hard spheres bouncing around and colliding with one another in a box of gas of a certain size under those conditions, whether the gas was oxygen, or carbon dioxide, or anything else.

The idea behind this is that in a box of gas there is mostly empty space, with the little hard balls whizzing about inside the box, colliding with one another and with the walls of the box. It doesn't matter what the little balls are made of – as far as the pressure on the walls of the box is concerned, all that matters is the speed of the particles and how often they hit it. The speed depends on the temperature (higher temperature corresponds to faster movement), and the number of hits per second depends on how many little hard balls there are in the box. So at the same temperature, pressure and volume, the number of particles must be the same.

This kind of model also explains the difference between gases, liquids and solids. In a gas, as we have said, there is mostly empty space, with the molecules hurtling through that space and colliding with one another. In a liquid there is no empty space, and the molecules can be envisaged as touching one another, but in constant movement, sliding past one another in an amorphous mass. In a solid the movement has all but stopped, and the molecules are locked in place, except for a relatively gentle jiggling, a kind of molecular running on the spot.

Avogadro’s idea wasn’t taken very seriously at the time (not even by Dalton). But at the end of the 1850s it was revived by Stanislao Cannizzaro, who realised that it provided a way of getting a measure of atomic and molecular weights. If you can find the number of molecules in a certain volume of one particular gas at a set temperature and pressure (the standard conditions are usually chosen as zero degrees Celsius and one standard atmospheric pressure), then you know the number of molecules present for any gas under those conditions. In order to find out how much each molecule weighs, you just have to weigh the gas and divide by that number.

For these standard conditions, you can choose a volume of gas which corresponds to two grams of hydrogen (two grams, not one, because each molecule of hydrogen contains two atoms, H_2). It works out as just over thirteen litres of gas. The number of molecules in such a volume is called Avogadro’s Number. The same volume of oxygen under the same conditions weighs thirty-two grams, and chemical evidence tells us that there are two oxygen atoms in each molecule. But it contains the same number of molecules as two grams of hydrogen. So we know, immediately, that one oxygen atom weighs sixteen times as much as one atom of hydrogen. This was a very useful way to determine relative atomic and molecular weights; but working out actual weights depends on knowing Avogadro’s Number itself, and that was harder to pin down.

There are several different ways to tackle the problem, but you can get some idea of the way it might be done from just one example, a variation on the theme used by Johann Loschmidt in the mid-1860s. Remember that in a gas there is a lot of empty space between the molecules, but in a liquid the molecules are touching one another. Loschmidt could calculate the pressure of gas in a container (under standard conditions) from Avogadro’s

Number, which determines the average distance travelled by the molecules between collisions (the so-called ‘mean free path’), and the fraction of the volume of the gas actually occupied by the molecules themselves. And he could find out how much empty space there was in the gas by liquefying it and measuring how much liquid was produced – or, indeed, by using measurements of the density of liquid oxygen and liquid nitrogen that other people had carried out. Since the particles are touching in a liquid, by subtracting the volume of the liquid from the volume of the gas he could find out how much empty space there was in the gas. So by adjusting the value of Avogadro’s Number in his pressure calculations to match the measured pressure, he could work out how many molecules were present.

Because the densities of liquid nitrogen and liquid oxygen used in his calculations were not as accurate as modern measurements, Loschmidt’s figure for Avogadro’s Number, derived in 1866, came out a little on the low side at 0.5×10^{23}. Using a different technique, Albert Einstein came up with a value of 6.6×10^{23} in 1911. The best modern value for the number is 6.022045×10^{23} or, in everyday language, just over six hundred thousand billion billion. This is the number of atoms in one gram of hydrogen, sixteen grams of oxygen, or in the gram equivalent of the atomic weight of any element. So each atom of hydrogen weighs 0.17×10^{-23} grams, and so on. Each molecule of air is a few hundred millionths of a centimetre across. At 0°C and one atmosphere pressure, one cubic centimetre of air contains 4.5×10^{19} molecules; the mean free path of a molecule of air is thirteen millionths of a metre, and an oxygen molecule in the air at that temperature travels at just over 461 metres per second (roughly 17,000 km per hour). So each molecule is involved in more than 3.5 billion collisions every second, producing the averaged-out feeling of a uniform pressure on your skin or on the walls of the room.

In fact, the kinetic theory of gases was first proposed by Daniel Bernoulli, as long ago as 1738. He was inspired by the work of Robert Boyle, in the middle of the seventeenth century. Boyle had discovered that when a gas is compressed (for example, by a piston) the volume of the gas changes in inverse proportion to the pressure – double the pressure and you halve the volume. Bernoulli explained this in terms of the kinetic theory, and also realised that the relationship between the temperature of a gas and its pressure (when you heat a gas its pressure increases, other things being equal) could also be explained in terms of the kinetic energy (energy of motion) of the little particles in the gas – heating the gas makes the particles move faster, so they have a greater impact on the walls of the container. But he was way ahead of his time. In those days most people who thought about heat at all thought that it was related to the presence of a kind of fluid, called caloric, which moved from one substance to another. Bernoulli's version of the kinetic theory made no impact at all on science at the time.

The kinetic theory was rediscovered twice (first by John Herapath in 1820, then by John Waterston in 1845), and ignored each time, before it finally became accepted by most scientists in the 1850s, largely as a result of the work of James Joule. A complete mathematical version of kinetic theory (a complete model) emerged in the 1860s, largely thanks to the work of Rudolf Clausius, James Clerk Maxwell, and Ludwig Boltzmann. Because this model deals with the averaged-out statistical behaviour of very large numbers of particles, which interact with one another like tiny billiard balls, bouncing around in accordance with Newton's laws of mechanics, it became known as statistical mechanics.

This is an impressive example of the way in which physical laws can be applied in circumstances quite different from those which were being investigated when the laws were discovered, and highlights the important difference between a law and a model. A law, like Newton's law of gravity, really is a universal truth. Newton discovered that every object in the Universe attracts every other object in the Universe with a force that is proportional to one over the square of the distance between the two objects. This is the famous 'inverse square law'. It applies, as Newton pointed out, to an apple falling from a tree and to the Moon in its orbit, in each case being tugged by the gravity of the Earth. It applies to the force holding the Earth in orbit around the Sun and to the force which is gradually slowing the present expansion of the Universe. But although the law is an absolute truth, Newton himself had no idea what caused it – he had no model of gravity.

Indeed, Newton specifically said in this context *hypotheses non fingo* ('I do not make hypotheses'), and did not try to explain why gravity obeyed an inverse square law. By contrast, Einstein's general theory of relativity provides a model which automatically produces an inverse square law of gravity. Rather than overturning Newton's ideas about gravity, as some popular accounts suggest, the general theory actually reinforces them, by providing a model to explain the law of gravity (it also goes beyond Newtonian ideas to describe the behaviour of gravity under extreme conditions; more of this later).

In order to be a good model, any model of gravity must, of course, 'predict' an inverse square law, but that doesn't mean that such a model is necessarily the last word, and physicists today confidently expect that one day they will develop a quantum theory of gravity that goes beyond Einstein's theory. If and when they do, though, we can be sure of one thing – that new model will still predict an inverse square law. After all, whatever new theories and models physicists come up with, the orbits of the planets around the Sun will still be the same, and apples won't suddenly start falling upwards out of trees.

Gravity, as it happens, is a very weak force, unless you have a lot of matter around. It takes the gravity of the entire Earth, pulling on an apple, to break the apple free from the tree and send it falling to the ground. But a child of two can pick the apple up from the ground, overcoming the pull of gravity. For atoms and molecules, rattling around in a box of gas, the gravitational forces between the particles are so tiny that they can be completely ignored. What matters here, as the nineteenth-century developers of statistical mechanics realised, are the other laws discovered by Newton – the laws of mechanics.

There are just three of these Newton's laws, which are so familiar today that they may seem like obvious common sense, but which form the foundations of all of physics. The first law says that any object stays still, or moves at a steady speed in a straight line (at constant velocity), unless it is pushed or pulled by a force. This isn't quite everyday common sense, because if you set something moving here on Earth (if I kick a ball, for example) it soon stops moving, because of friction. Newton's insight was to appreciate how things behave when there is no friction – things like rocks moving through space or, indeed, atoms whizzing about in a box of gas (incidentally, although he never developed a kinetic theory of gases, Newton was himself a supporter of the atomic model, and wrote of matter being made up of 'primitive Particles … incomparably harder than any porous bodies compounded of them; even so very hard, as never to wear out or break in pieces').

Newton's second law says that when a force is applied to an object it accelerates, and keeps on accelerating as long as the force is applied (acceleration means a change in the speed of an object, or the direction in which it is moving or both; so the Moon is accelerating around the Earth, even though its speed stays much the same, because its direction is constantly changing). The acceleration produced by the force depends on the strength of the force divided by the mass of the object (turning this around, physicists often say that the force is equal to the mass times the acceleration). This does match up with common sense – it is harder to push objects around if they have more mass. And Newton's third law says that when one object exerts a force on another there is an equal and opposite reaction back on the first object. When I kick a ball (or, if I am foolish enough to do so, a rock), the force my foot exerts on the ball (or rock) makes it move, and the equal and opposite force the object exerts on my foot can be clearly felt.

More subtly, just as the Earth tugs on the Moon through gravity, so there is an equal and opposite force tugging on the Earth. Rather than the Moon orbiting the Earth, we should really say that they each orbit around their mutual centre of gravity – but the Earth is so much more massive than the Moon that, as it happens, this point of balance actually lies below the surface of the Earth. Strictly speaking, too, the equality of action and reaction means that when the apple is falling to the ground the whole Earth, tugged by the apple, moves an infinitesimal amount 'up' to meet the apple. It is Newton's third law that explains the recoil of a gun when it is fired and the way a rocket works by throwing matter out in one direction and recoiling in the opposite direction.

These three laws apply to the Universe at large, which is where Newton applied them to explain the orbits of the planets, and to the everyday world, where they can be investigated by doing experiments such as rolling balls down inclined planes and measuring their speed, or by bouncing balls off one another. But, because they are, indeed, universal laws they also apply on the scale of atoms and molecules, providing the mechanics which, as I have mentioned, is the basis of statistical mechanics and the modern kinetic theory of gases – even though the modern kinetic theory was developed nearly two centuries after Newton came up with his three laws of mechanics, and he had never applied his laws in this way. Newton did not really invent the laws at all – they are laws of the Universe, and they operated in the same way before he wrote them down, just as they operate in places he did not happen to investigate.

There are good reasons why the kinetic theory, and statistical mechanics, were at last taken on board by scientists in the middle of the nineteenth century. By that time, the ground had been prepared by the study of thermodynamics (literally, heat and motion), which was of immense practical importance in the days when the industrial revolution in Europe was being driven by steam power.

The principles of thermodynamics can also be summed up in three laws, and they have a wide-ranging importance which applies across science (indeed, across the Universe), not just to the design and construction of efficient steam engines. The first law of thermodynamics is also known as the law of conservation of energy, and says that the total energy of a closed system stays the same. The Sun is not a closed system, and is pouring energy out into space; the Earth is not a closed system, and it receives energy from the Sun. But in a process like a chemical reaction taking place in an insulated test tube, or in the processes involved in statistical mechanics where little hard particles bounce around inside a box, the total amount of energy is fixed. If one fast-moving particle carrying a lot of kinetic energy collides with a slow-moving particle which has little kinetic energy, the first particle will probably lose energy and the second particle will gain energy. But the total energy carried by the two particles before and after the collision will be the same.

Since Einstein came up with the special theory of relativity early in the twentieth century, we know that mass is a form of energy, and that under the right circumstances (such as inside a nuclear power station, or at the heart of the Sun) energy and mass can be interchanged. So today the first law of thermodynamics is called the law of conservation of mass-energy, not just the law of conservation of energy.

The second law of thermodynamics is arguably the most important law in the whole of science. It is the law which says that things wear out. In terms of heat – the way the law was discovered in the days of steam engines – the second law says that heat will not flow from a colder place to a hotter place of its own volition. If you put an ice cube in a cup of hot tea, the ice melts and the tea gets cooler – you never see a cup of lukewarm tea in which the tea gets hotter while an ice cube forms in the middle of the cup, even though such a process would not violate the law of conservation of energy. Another manifestation of the second law is the way that a brick wall, left undisturbed, will be worn down and crumble away, while a pile of bricks, left undisturbed, will never assemble themselves into a brick wall.

In the 1920s the astrophysicist Arthur Eddington, slightly tongue-in-cheek, summed up the importance of the second law in his book *The Nature of the Physical World:*

> 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.

The second law is also related to the concept known as entropy, which measures the amount of disorder in the Universe, or in a closed part of the Universe (such as a sealed test tube in the laboratory). Entropy in a closed system cannot decrease, so any change in the system moves it towards a state of higher entropy. The 'system' of an ice cube floating in a cup of tea has more order (less entropy) than a cup of lukewarm tea, which is why the system shifts from the ordered state to the disordered state.

The Universe as a whole is a closed system, so the entropy of properties are related to the number of protons alone – more of this in the next chapter – but neither the proton nor the neutron were known in the nineteenth century so there was no way that Mendeleyev could have explained the physical basis for this slight reordering of the elements in a table based on atomic weights, and he relied on the chemical evidence of similarities.

Mendeleyev's boldest step, and the one which eventually led to the widespread acceptance of his periodic table as something related to the fundamental properties of the elements, (and not an arbitrary convention like the alphabet) was his willingness to leave gaps in the table where there was no known element with properties that 'belonged' in a certain place. By 1871 Mendeleyev had produced a table containing the sixty-three elements known at the time, showing the striking periodicity in which families of elements with atomic weights that are multiples of eight times the atomic weight of hydrogen have similar chemical properties to one another. But to make the pattern work, even after minor adjustments like swapping the positions of tellurium and iodine, he had to leave three gaps in the table, and he boldly predicted that new elements would be found with properties (which he specified) corresponding to the places of those gaps in his table. The three elements, with exactly the predicted properties, were discovered over the next fifteen years: gallium, in 1875; scandium, in 1879; and germanium, in 1886.

In the classic tradition of science ('if it disagrees with experiment then it is wrong') Mendeleyev had made a prediction, and it had been proved correct. This persuaded people that the periodic table was important, and as more new elements were discovered and each one was found to fit into Mendeleyev's table, acceptance of his ideas turned to enthusiasm. There are now ninety-two elements known to occur naturally on Earth, and more than twenty heavier elements which have been created artificially in particle accelerators. All of them fit into Mendeleyev's table, allowing for some improvements to the layout of the table which have been made in the twentieth century to take account of our modern understanding of the structure of atoms.

But even the success of Mendeleyev's periodic table in the last third of the nineteenth century still did not persuade everyone of the reality of atoms. The final acceptance of the 'atomic

Copyright

A Weidenfeld & Nicolson ebook

First published in Great Britain in 1998 by Weidenfeld & Nicolson
First published in ebook in 2015 by Weidenfeld & Nicolson

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

ISBN: 978 1 4746 0273 0

The Orion Publishing Group Ltd
Carmelite House
50 Victoria Embankment
London, EC4Y 0DZ

An Hachette UK company

www.orionbooks.co.uk