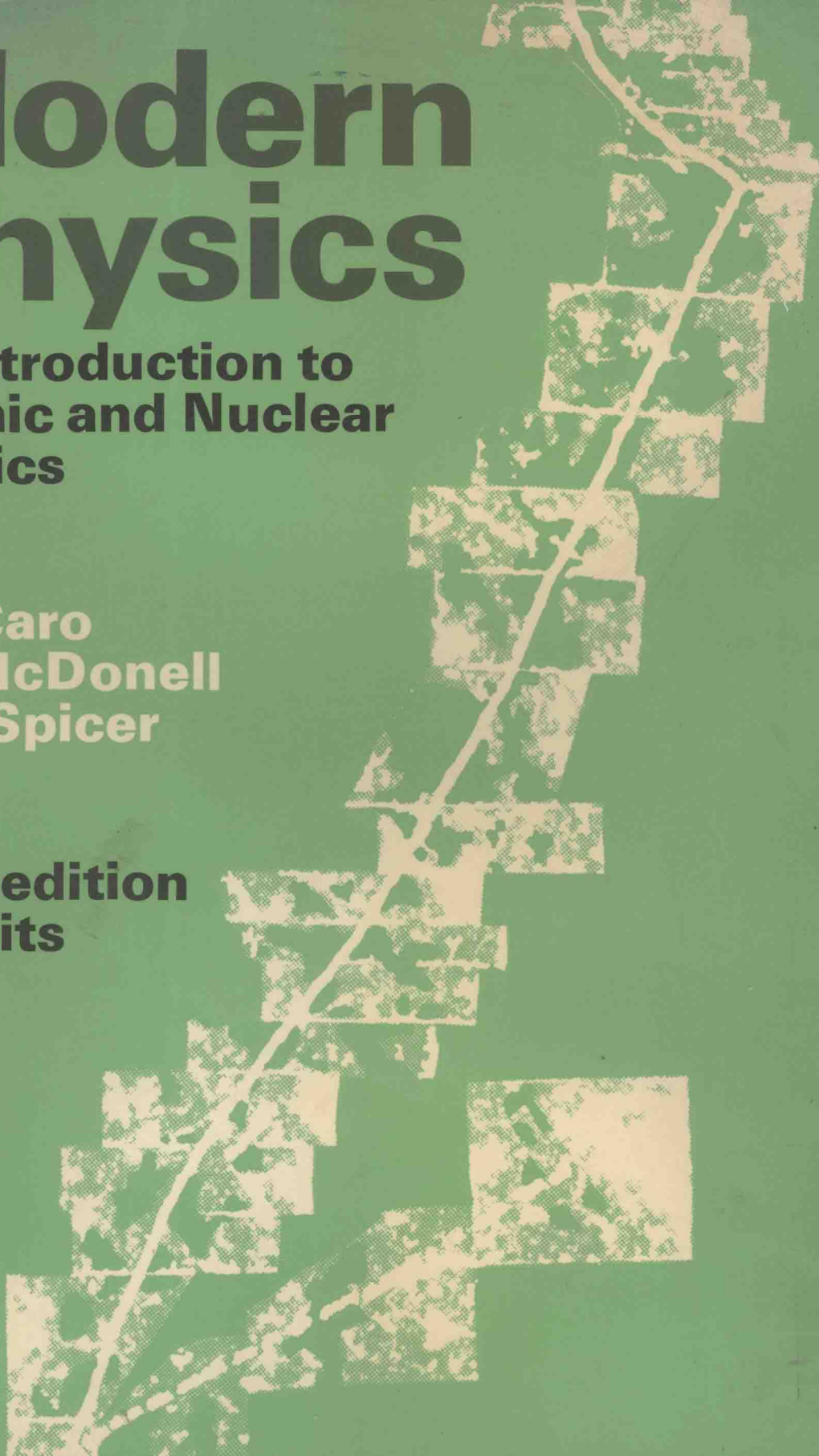


Modern Physics

**An introduction to
Atomic and Nuclear
Physics**

**D E Caro
J A McDonnell
B M Spicer**

**third edition
SI units**



modern physics

An Introduction to Atomic and Nuclear Physics

third edition

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Edward Arnold

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First published 1961 by
F. W. Cheshire Publishing Pty Ltd
First published in Great Britain 1962 by
Edward Arnold (Publishers) Limited
41 Bedford Square, London WC1B 3DP
Reprinted 1964, 1965, 1966
Second edition 1971
Reprinted 1973, 1976
Third edition 1978

ISBN 0 7131 2704 X

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Text set in 11/12 pt IBM Press Roman, printed by photolithography,
and bound in Great Britain at The Pitman Press, Bath

Preface to the third edition

In the seventeen years since the first edition of this book appeared there have been some notable advances in this branch of physics. As a result, several of its chapters have been in need of updating – and this is even true of some which had already been modified, in the light of new knowledge, in the second edition of 1971. The outstanding example is in the field of fundamental particles (Chapter 17). New information obtained from modern accelerators has led to the identification of a formidable array of particles and a completely new theoretical structure has emerged to accommodate them. Although no satisfactory account of the detailed arguments in support of this structure can be given in a book of this kind we have thought it advisable to attempt a description of its main features. This section has been revised, for this edition, by Mr. Lindsay J. Martin, whose assistance is gratefully acknowledged.

Advances in techniques have been included and some material which, with the passage of time has assumed relatively less importance, has been dropped – particularly in the revision of Chapters 11 and 13. We have attempted to improve upon the presentation of the material by some rewriting and reordering in several places, principally in Chapters 3, 4, 9 and 14.

As in the second edition, SI units have been adopted throughout, despite our having some reservations about their appropriateness in one or two cases. (Is the becquerel really a convenient unit for the activity of a radioactive source?) As before, the only notable departure from this practice is the retention of the electron-volt as the appropriate unit of energy in many parts of the text.

We have received helpful advice from a number of users of the book and would like particularly to thank those who provided references for extensions to the Bibliography in Appendix A. In many places we have adopted, with gratitude, suggestions made by our correspondents as to matters of content and presentation. In others, we have preferred to retain the original style of the book. No textbook can be entirely to the liking of all its users – changes which appeal to some will not necessarily be attractive to others; and in considering proposed changes, we have, if in doubt, fallen back to the idiosyncratic position of adopting the version which appeals to the authors.

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Preface

Some years ago it was quite common for many University students to receive their introduction to physics in the first year of their University course. This has now become quite exceptional in the University of Melbourne, where we find that nearly all students taking first-year physics – and this includes those doing such courses as Medicine, Dentistry and Agricultural Science – have had one or two years' physics at school. This situation is, we believe, common to many Universities. This being the case, there is room, in such first-year courses, for quite an appreciable amount of atomic and nuclear physics, and it is in an attempt to meet the specific needs of these courses that this book has been written.

The standard aimed at has been deliberately made higher than one would normally expect a student to reach after about three years' study of physics as a separate subject. Thus the book should provide a considerable amount of reading material extending beyond the immediate requirements of most conceivable introductory courses in atomic and nuclear physics.

While the material covered has been determined principally by the needs of first-year University students, we have tried to make the book useful and readable for students in their last school year. Some chapters, particularly in the early part of the book, should be completely within the range of such students. In nearly all other chapters, the development is such that these students should be able to follow the early sections without any difficulty and without need to refer back to more difficult sections of preceding chapters. Certainly the subject matter goes, in matters of detail, a long way further than any syllabus which we envisage as being appropriate to a secondary school course in physics. However, a good student embarking on such a course might, we hope, be encouraged to explore the subject beyond his immediate requirements for examination purposes.

In the belief that students should understand *how* a theory was developed as well as *what* the theory is, we have adopted a historical approach to the material presented. Nuclear and atomic physics is a living subject and a study of its growth can be an exciting experience. Furthermore, in studying the way new concepts evolve, the student gains an understanding of the methods employed by scientists to unravel the problems confronting them. We are aware that some students resent being taught 'history' in a physics course. However, we believe that most of them enjoy the subject if it is taught as a 'detective story', showing the way in which contributions from apparently unrelated investigations are brought together to build up a coherent picture of a new field of science.

The standard of presentation varies quite considerably – and intentionally – throughout the book. For example, in Chapter 9 we have gone further in discussing the electron structure of atoms than is usual in texts of this nature. We have done so because the modern teaching of chemistry demands a relatively detailed knowledge of electron states. It seems to us to be questionable that students should be asked to use this material before some justification has been provided for its physical basis. Again, in the chapters devoted to nuclear reactions and nuclear structure, we have provided more material than is customary in an introductory course. It has been our experience that the good student wishes to learn something about nuclear structure at an early stage.

We have not hesitated to use calculus when it seemed to provide the easiest method of deriving a particular relation. However, there are few places in the book where it appears and these derivations can be omitted, if necessary, without weakening the discussion at any point. In most cases an elementary knowledge of algebra is sufficient. It has been assumed that the reader has some knowledge of the elements of mechanics, electricity and magnetism, but no deep experience is demanded. The MKSA system of units has been used throughout.

In the text we have provided references to original papers describing significant experiments and new theoretical concepts. Exceptions have been made where a new theoretical idea has been couched in mathematical language which is too sophisticated for the young student.

We would like to acknowledge permission to publish material drawn from many sources. Individual acknowledgments have been made at appropriate points elsewhere in the book. Our first-year students have unwittingly played no small part in the production of this book. We have taught from it and the reactions of students have helped us to appreciate their problems, so that

perhaps the path of future students may be strewn with rather fewer difficulties.

Finally we wish to express our appreciation of the assistance of Mrs. P. George who has been able to decipher the manuscript and from it prepare an excellent typescript. Our thanks are due to Mr. L. McBride, who has produced the line drawings, to Mr. J. Smith, for assistance in the reproduction of photographs, to Miss J. Filshie, for considerable secretarial assistance, and to our publishers, who have shown great patience and co-operation.

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Contents

The Atomic Theory of Matter

1 The Growth of Atomic Concepts

1.1	Early Ideas about the Structure of Matter	2
1.2	The Rise of Atomic Theory in Chemistry	5
1.3	Physical Evidence for the Atomic Theory	5

The Turn of the Century: Some Momentous Advances in Physics

2 The Identification of the Electron

2.1	The Conduction of Electricity in Gases	9
2.2	Early Observations of Cathode Rays	10
2.3	The Charge and Mass of Cathode Rays	11
2.4	The Electron	13
2.5	The Accurate Measurement of the Charge of the Electron	15
2.6	The Accurate Measurement of the Specific Charge of the Electron	17

3 The Photoelectric Effect

3.1	The Early Observations of Hertz and Others	20
3.2	Theoretical Difficulties	21
3.3	Einstein's Predictions and the Quantum Theory	22
3.4	The Electron-volt	23
3.5	The Experimental Verification of Einstein's Predictions	24
3.6	Variation of Photoelectric Emission with Intensity	25

4 X-Rays

4.1	The Discovery of X-rays	28
4.2	Röntgen's Preliminary Communication	28
4.3	Röntgen's Second and Third Papers	29
4.4	Summary of the Properties of X-rays	30
4.5	The Coolidge Tube	31
4.6	The Detection of X-rays	32
4.7	Diffraction of X-rays by Crystals	33
4.8	The Bragg X-ray Spectrometer	34
4.9	X-ray Spectra	36
4.10	The Continuous Spectrum	37
4.11	The Absorption of X-rays	38
4.12	X-ray Crystallography	39
4.13	Radiation — Waves or Particles?	40

5 The Discovery of Radioactivity

5.1	Becquerel's Experiments	43
5.2	The First Radiochemists	44
5.3	The Identification of the Three Types of Radiation	46
5.4	The α -rays	47
5.5	The β -rays	51
5.6	The γ -rays	53
5.7	Radioactive Transformations	53

6 Relativity

6.1	Observers and their Frameworks	58
6.2	The Propagation of Light and the Aether	59

6.3	Frameworks and Transformations	61
6.4	Consequences of the Lorentz Transformation	63
6.5	Energy and Momentum	64
6.6	Particles and Photons	66

The Structure of the Atom Emerges

7 The Radiations Given off by Atoms

7.1	The Early Days of Spectroscopy	69
7.2	The Search for Regularities in Line Spectra	70
7.3	The Main Features of Atomic Spectra	72
7.4	Additional Series in the Hydrogen Spectrum	73

8 The Electric Charges in Atoms

8.1	Early Ideas about the Structure of the Atom	75
8.2	The Nuclear Atom: Preliminary Experiments	75
8.3	The Nuclear Atom: Rutherford's Theory	76
8.4	The Nuclear Atom: Test of Rutherford's Theory	78
8.5	The Number of Electrons in an Atom	80

9 The Development of the Nuclear Atom

9.1	Objections to Rutherford's Model of the Atom	82
9.2	Bohr's Radical Postulates	82
9.3	Application of Bohr's Postulates to the Hydrogen Atom	83
9.4	First Refinement of Bohr's Theory	85
9.5	Further Predictions of the Bohr Theory	87
9.6	Energy Levels of the Hydrogen Atom: Quantum Numbers	88
9.7	The Experiments of Franck and Hertz	88
9.8	Elaboration of Bohr's Theory	89
9.9	Matter Waves	90
9.10	Stationary States and Quantum Numbers	92
9.11	The Periodic Table: Pauli's Exclusion Principle	94
9.12	Characteristic X-rays: Moseley's Law	95
9.13	The Structure of the Atom	98
9.14	Wave Mechanics	99

10 The Masses of Atoms

10.1	The Discovery of Positive Rays	103
10.2	Thomson's Experiments	103
10.3	The Discovery of Isotopes	105
10.4	The Work of Aston	106
10.5	Dempster's Mass Spectrometer	108
10.6	Isotopic and Atomic Masses	109

The Nucleus Comes Under Fire

11 Detectors of Nuclear Radiations

11.1	Introduction	112
11.2	The Ionization Chamber	112
11.3	The Geiger-Müller Counter	114
11.4	The Scintillation Counter	116
11.5	Visual Observations of the Tracks of Charged Particles	118
11.6	The Photographic Emulsion	120

11.7	Spectroscopy of Charged Particles	120
11.8	The Detection of Neutrons	122
12 The First Shots are Fired with Alpha-particles		
12.1	Collisions of α -particles with Hydrogen	123
12.2	The Discovery of Artificial Transmutation	123
12.3	Interpretation of this Disintegration Phenomenon	126
12.4	Conservation of Energy in Nuclear Reactions	127
12.5	The Discovery of a Strange Type of Radiation	128
12.6	Evidence for the Existence of the Neutron	129
13 Particle Accelerators: The Development of Nuclear Artillery		
13.1	Introduction	133
13.2	Electrostatic Accelerators	133
13.3	The Van de Graaff Generator	135
13.4	The Tandem Accelerator Principle	136
13.5	The Linear Accelerator	136
13.6	The Cyclotron	138
13.7	The Synchrocyclotron and Synchrotron	141
13.8	The Betatron	142
14 Nuclear Reactions		
14.1	Introduction	146
14.2	Cockcroft and Walton's Experiment	146
14.3	Nuclear Structure and Conservation Laws	147
14.4	The Q -value of a Nuclear Reaction	148
14.5	A Picture of Nuclear Reactions	149
14.6	Nomenclature of Nuclear Reactions: Proton-induced Reactions	150
14.7	Deuteron-induced and Alpha-induced Reactions	150
14.8	Photon-induced Reactions	151
14.9	Neutron-induced Reactions	152
14.10	Probability of a Nuclear Reaction: Cross-sections	152
14.11	Details of Reaction Processes	153
15 Radioactivity and Nuclear Structure		
15.1	Artificial Radioactivity	156
15.2	Alpha-radioactivity	157
15.3	The Beta-disintegration Process	159
15.4	Radioactive Series	161
15.5	Nuclear Binding Energy	162
15.6	Nuclear Stability	166
15.7	Nuclear Forces	168
16 Nuclear Energy		
16.1	The Possibility of Obtaining Energy from the Nucleus	169
16.2	The Discovery of Fission	169
16.3	The Fission Reaction	170
16.4	The Chain Reaction in Uranium	171
16.5	The Thermal Reactor	172
16.6	The 'Atomic' Bomb	174
16.7	Fast Reactors and Breeding	175
16.8	Fusion and the Future	176

17 More Fundamental Particles Are Discovered	
17.1 Cosmic Radiation	178
17.2 The Positron	180
17.3 Creation and Annihilation of Electron-positron Pairs	182
17.4 The μ -meson	183
17.5 The π -meson	184
17.6 Particles and Anti-particles	186
17.7 The Fundamental Forces	186
17.8 Leptons	186
17.9 Strongly Interacting Particles	187
17.10 Baryons	189
17.11 Mesons	189
17.12 Quarks	190
Appendix A: Bibliography	192
Appendix B: Table of Masses and Abundances of Naturally Occurring Isotopes	197
Appendix C: Table of Naturally Occurring Radioactive Isotopes	203
Appendix D: Table of Physical Constants	207
Index	208

The Atomic Theory of Matter

Most of this book is concerned with the behaviour of atoms and the particles of which they are composed. Nobody has ever seen, or is ever likely to see an atom, if by 'seeing' we mean the usual methods of visual observation, assisted perhaps by optical magnification. But if 'seeing' atoms is taken to mean the observation of effects which can readily be explained in terms of an atomic theory of matter and are difficult to account for in any other way, then seeing atoms and sub-atomic particles has become almost a commonplace experience for a great many people. Look at the luminous figures on a clock through a magnifying glass — in the dark, the faint glow of light will then appear as a host of tiny flashes. Listen to a Geiger counter of the type used in geological surveying — each click records a separate event. These flashes and clicks are each the result of the detection, or 'seeing' of an individual particle from the atomic world.

But does this atomic world really exist? Is it reasonable to regard atoms as actual physical objects even though they can never really be seen or handled? These are very reasonable questions to ask and until about 1910 there were still some eminent scientists who found themselves unhappy about accepting the atomic theory as giving a true picture of the nature of matter. They felt that a theory based on such completely intangible entities, as atoms then appeared to be, was hardly acceptable as a really firm basis on which to build a complete picture of the universe. With the passage of time, however, the atomic theory has come to be universally accepted. Such an enormous variety of phenomena, ranging from the evolution of stars to the structure of the minute constituents of living organisms and including such vast fields as chemistry and electronics are capable of description in terms of the behaviour of atoms and their components. It is hardly surprising that we are now quite happy to believe in the reality of atoms.

1 The Growth of Atomic Concepts

1.1 Early Ideas about the Structure of Matter

The Greek philosophers, Democritus and Leucippus, about 2400 years ago, were the first to suggest in their writings that matter might not be infinitely divisible, but might in fact consist ultimately of particles which cannot be further subdivided. From the Greek word *atomos*, meaning 'indivisible', they proposed the name *atom* for these 'fundamental' particles. Less than a century later, however, Aristotle was at the height of his powers and was proposing an entirely different picture of the constituents of matter. In his view, all matter in the world in which we live is made up of mixtures, in various proportions, of the four elements, Fire, Air, Water and Earth. Each of these elements was supposed to have its 'natural place' and 'natural movement', Fire having the greatest tendency to rise and Earth having the greatest tendency to fall. In this way one could readily explain, for example, why gases, being largely composed of the element Air, rise through liquids; while solids, in which Earth is the predominant element, fall through both liquids and gases. In addition, Aristotle proposed a fifth element, the Quintessence, from which all heavenly bodies are formed, this element having its natural place in circular orbits around the earth.

Now while this five-element theory may seem fanciful to us today, it did provide an explanation of sorts for a large number of phenomena and this, after all, is one criterion by which any scientific theory must be judged. Science, at the time, was a subject for philosophy rather than for experimentation, and verification of the few quantitative predictions made by Aristotle was largely neglected. The prestige of Aristotle, who had made important contributions to many fields of learning, resulted in the acceptance of his views as authoritative for about 2000 years. Various authors had questioned these theories during this period, but no one had presented arguments of sufficient weight to overthrow them until Galileo was able to do so. He, by consistent use of experimental investigation, showed that Aristotle's pronouncements on the laws governing the motion of falling bodies were untenable and that, as Copernicus and Kepler had proposed, the earth was in motion around the sun, in contradiction to the

Aristotelian view that the earth constituted the centre of the universe. The way was now cleared for a revision of other physical theories, in particular the 'five-element' picture of the structure of matter.

1.2 The Rise of Atomic Theory in Chemistry

One of the first steps forward was taken by the Englishman, Robert Boyle, the 'Father of Chemistry'. In his book, *The Sceptical Chymist*, which appeared in 1661, he introduced for the first time the concept of chemical elements as distinct from mixtures and compounds. Boyle put forward this idea with considerable clarity:

And to prevent mistakes I must advertize you, that I now mean by elements, . . . certain primitive and simple, or perfectly unmingled bodies; which not being made of any other bodies, or of one another, are the ingredients of which all those call'd perfectly mixed bodies are immediately compounded, and into which they are ultimately resolved: . . .

In the light of these principles the identification of elements and the understanding of the ways in which they combine to produce chemical compounds was carried forward with considerable success by the chemists of the eighteenth century, outstanding amongst whom were such men as Joseph Black, the Scottish physician, the Yorkshireman Joseph Priestley and the greatest of them all, the French chemist, Antoine Lavoisier. From this study of chemical compounds, two significant facts emerged. Lavoisier was careful to measure the amounts of the reacting substances which took part in various reactions and, generalizing his results, proposed as ' . . . an incontestable axiom . . . ' the principle that *the total mass of an isolated reacting system remains constant*. In addition, Proust put forward his *law of definite proportions*, in which he pointed out that *the proportions by weight of the elements present in any chemical compound are entirely independent of the manner in which the compound is prepared*. With these laws and the concepts of the existence of elements and of their combinations with one another firmly established, the stage was set for the appearance of the atomic theory, as proposed by the Englishman, John Dalton.

The main points of Dalton's theory, as expounded in his book *A New System of Chemical Philosophy*, which was completed in 1810, may be summarized as follows:

(1) *Matter consists of indivisible particles, which Dalton referred to as atoms, and each element consists of identical atoms of a type characteristic of that particular element. Atoms themselves are completely unchangeable.*

These concepts were entirely consistent with the established picture of elements as substances which were the basic components of all chemical compounds and which could not be transformed from one to another.

(2) *When elements combine to form compounds, their atoms are not created or destroyed, but group together in simple combinations which constitute the smallest component particles (or, as we now call them, molecules) of these compounds. Then, reactions between compounds or between elements and compounds simply result in the rearrangement of the atoms into new molecules.*

This description of the fundamental processes occurring in chemical combinations clearly gave a satisfactory explanation of Lavoisier's law of conservation of mass and of Proust's law of definite proportions.

Dalton also put forward two rules governing the possible ways in which compounds could be formed. In an endeavour to give some concrete form to his ideas as to how atoms combined into molecules he suggested his *rule of simplicity* which states that, *when only one combination of two elements can be obtained it should be assumed to be a simple binary one*, that is, one atom of each element combining to form a molecule of two atoms. The next simplest combination, if a second compound of the same two elements is known, would be a three-atom molecule.

The other of Dalton's laws was his *law of multiple proportions*, in which he proposed that, *if two elements A and B combine to form various compounds and if the mass of element A in each compound is kept constant, then the masses of element B present in these compounds will be in the ratios of small integers*. This law was well substantiated by a mass of information which had been accumulated before Dalton's time, but which had

never been analysed in such a way as to bring out this particularly simple relationship. It is, of course, in complete accord with the concept of atoms combining in fixed proportions for every compound.

Perhaps Dalton's most important contribution was towards the determination of the relative masses of atoms. Although many of the results he obtained are incorrect, he must be credited with the enormous step forward of assigning some measurable quantity to an atom — its mass — and using this quantity as a reliable guide in the identification of elements and the analysis of compounds. In order to find the relative atomic masses of two elements A and B, the first step is to determine their proportions in a compound. If the masses of the atoms of A and B are m_A and m_B respectively, and if there are a atoms of A and b atoms of B in each molecule of the compound, then

$$\begin{aligned} R &= \frac{\text{Mass of element A in the compound}}{\text{Mass of element B in the compound}} \\ &= \frac{\text{Mass of the atoms of A in the molecule}}{\text{Mass of the atoms of B in the molecule}} \\ &= \frac{a m_A}{b m_B}. \end{aligned}$$

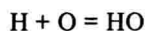
Now the ratio R is determined experimentally, but the ratio $\frac{a}{b}$ must, at this stage, be assumed; and this is the point at which Dalton used his rule of simplicity. By calculations of this sort, extended to many compounds, Dalton was able to draw up a list of what he considered to be the most reasonable values of atomic masses. These he referred to the lightest atom, hydrogen, whose mass he took to be his unit of atomic mass.

However, there was one step in Dalton's argument which proved to be untenable, namely his rule of simplicity. In the case of water, for example, he assumed that the molecule consisted of one atom of hydrogen together with one of oxygen, so that $\frac{a}{b} = 1$, and the oxygen atom is then taken to be 8 times heavier than the hydrogen atom, since oxygen and hydrogen combine in the ratio 8 : 1 by mass to form water. Despite errors of this kind, Dalton must be given the credit for compiling the first table of atomic masses and

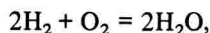
for drawing attention to the importance of these quantities.

At almost the same time, the French chemist Joseph Gay-Lussac, an exceedingly careful and capable experimenter, showed that, in reactions between gases under conditions of constant temperature and pressure, *the volumes of the reacting gases and of the product gas were always expressible in simple ratios*. Such a result once again suggested, qualitatively, a simple combination between atoms and molecules, in accord with Dalton's fundamental ideas. But in matters of detail, Dalton found it hard to accept Gay-Lussac's result. For example, he was unable to reconcile the combination of two volumes of hydrogen with one of oxygen to form two volumes of water vapour with his own rule of simplicity, in which one atom of hydrogen combined with one of oxygen to form one water vapour molecule. Furthermore, since he believed that different atoms occupied different volumes, he was unable to understand how, in other cases, the quantities of the reacting gases which combined, and which must then have had equal numbers of atoms, could have equal volumes.

This conflict was resolved by the Italian physicist Amedeo Avogadro who, guided by Gay-Lussac's law, put forward the hypothesis that *the number of molecules in a given volume of a gas is the same for all gases*. Then if, in a particular case, one volume of gas A combined with three volumes of gas B to form two volumes of gas C, in accordance with Gay-Lussac's law of simple ratios, Avogadro's interpretation of such a reaction would be that each molecule of gas A was combining with three molecules of gas B to form two molecules of gas C. This, of course, implied that the molecules were divisible, each consisting of more than one atom. Let us consider once again the combination of hydrogen and oxygen to form water vapour. In terms of our modern notation, Avogadro replaced the concept of hydrogen and oxygen combining according to the equation



with the proposition that the reaction was more correctly described by



thus retaining all of the important features of Dalton's theory and incorporating the results of Gay-Lussac's precise measurements.

Unfortunately, Avogadro's paper in which he set forth these ideas was, in places, rather obscurely worded; in particular, the distinction in meaning between the terms 'integral molecule' and 'solitary elementary molecule', which he used to describe the particles which we now refer to as 'molecules' and 'atoms' respectively, was not always as clear as it might have been. Partly for this reason and partly because of the esteem in which Dalton's work was held, Avogadro's paper was largely ignored. In addition, there was one particularly weighty objection which was quite validly raised. If two similar atoms tend to stick together to form a molecule, why do not *all* of the similar atoms in a gas cling together, so that the gas condenses? A century was to pass before a satisfactory answer to this question emerged.

As a result of this neglect of Avogadro's ideas, attempts to produce a consistent set of atomic masses — the values of which, as we have seen, depend on a knowledge of the atomic composition of molecules — resulted in confusion which worsened and persisted for a period of nearly fifty years. Ultimately, however, in 1858 Cannizzaro, a fellow-countryman of Avogadro, was responsible for 'reviving' Avogadro's paper and clarifying once and for all the relationship between molecule and atom.

Once this stage had been reached, the determination of atomic and molecular masses could be carried out in an orderly manner. It soon became apparent that, if the scale of atomic masses was based on that of hydrogen being taken as unity, all atomic masses were almost integers. However, if oxygen was chosen instead as the standard and its atomic mass assigned to be exactly 16, the atomic masses of the other elements came, on the average, to be closer still to integers. So this scale was adopted. Now if the molecular masses of two gases are m_A and m_B , equal volumes of these two gases, since they contain equal numbers of molecules, will have masses in the ratio $m_A:m_B$. Conversely, amounts of these two gases with masses m_A kg and m_B kg respectively will occupy the same volume and contain the same number of molecules. In

general then, the number of molecules in m kg of a gas whose molecular mass is m is a constant. By convention we usually deal with a smaller constant — the number of molecules in $10^{-3} m$ kg (or m gram) of the gas. This quantity of the gas is called a *mole*. This constant, for which we use the symbol L , is known as the Avogadro constant and turns out to be of fundamental importance in physics. If it can be determined, the masses of atoms can then be measured in kilograms — i.e. the convenient unit of the 'atomic mass scale' can be related to the kilogram. The molecular mass of oxygen is 32 so that there are L oxygen molecules or $2L$ oxygen atoms in 32 gram of that gas. Thus the mass of the oxygen atom is just $\frac{16}{L} \times 10^{-3}$ kg and the masses of other atoms are similarly determined — hence the significance of the measurement of L . It should be noted that while L has been defined in terms of a discussion of molecules, it is equally true that there are L atoms of any element of atomic mass m in $10^{-3} m$ kg (i.e. one mole) of that element.

L is in fact an exceedingly large number — 6.02×10^{23} entities per mole. The first determination of the Avogadro constant with any claim to precision was not made until almost 100 years after the appearance of Avogadro's paper and the measurement was based, not on any chemical properties at all, but on physical phenomena described by the *Kinetic Theory of gases*.

1.3 Physical Evidence for the Atomic Theory

The Swiss physicist Daniel Bernoulli was apparently the first to suggest, in 1738, that a gas should be thought of as consisting of a vast number of particles, all in rapid motion. In his view, these particles should be regarded as perfectly elastic, so that no energy is lost by collisions between them. Then the pressure on the walls of the containing vessel would be provided by the enormous number of impacts of these tiny particles, a process which would go on continuously. Bernoulli went so far as to show that this model would result in the product of the pressure and volume of a gas at constant temperature being constant, in accordance with Boyle's law. Strangely enough this was another case of an important idea

which was not followed up for a long time.

It was not until 1847 that the Kinetic Theory was brought back into prominence by James Joule, who had studied under Dalton. In the meantime, however, there had been several developments which very much strengthened its basic assumptions. In the first place the chemists, led by Dalton, had provided solid support for the atomicity of matter. Then there was the important contribution by Count Rumford, confirmed by a whole series of experiments by Joule himself, in which heat and motion were shown to be only different forms of the same quantity — energy. Lastly, there was the law proposed by the German doctor, Julius Mayer, which has become one of the foundations of physical thought — the *law of conservation of energy*. Thus the ground was well prepared for the acceptance of the Kinetic Theory of gases, which was to be remarkably successful in accounting for a wide range of physical phenomena.

The principal assumptions of the Kinetic Theory as put forward by Joule are:

- (1) *A gas consists of a large number of molecules, all of which are in rapid and random motion.*
- (2) *These molecules are very small compared with the average distance between them, and have no influence on one another, except when they happen to collide.*
- (3) *Collisions between molecules are perfectly elastic; that is, no energy losses occur when such collisions take place.*
- (4) *The temperature of the gas is directly proportional to the average kinetic energy of its molecules; thus, when a gas is heated at constant volume, the energies of its molecules are increased and its temperature rises.*

We shall not describe in any detail the way in which these assumptions (and further refinements which were added by Maxwell, Boltzmann and others in the latter half of the nineteenth century) can be used to derive a remarkable range of quantitative results about the behaviour of gases. It is sufficient for our purpose to realize that the theory is able to predict such apparently diverse results as the laws of Boyle and Charles which govern the relationships between the pressure, volume and temperature of a gas, the absence of an atmosphere around the moon, the fact that

the viscosity of a gas is largely independent of its pressure and density, the rates at which gases diffuse through one another, the specific heats of gases and many other experimentally verifiable relationships between physical quantities.¹

These impressive successes of the Kinetic Theory, together with the rapid and consistent growth of chemistry and chemical theory appeared to provide overwhelming evidence for the basic assumption in each case — the existence of atoms and molecules. Nevertheless there remained a few quite eminent sceptics who were unable to believe in the reality of particles which no one had ever seen. They felt that *direct* evidence for the existence of atoms was still lacking. These final doubts were removed in 1912 by the experiments of the French chemist Jean Perrin. Perrin undertook the first quantitative study of a phenomenon which had been discovered by the botanist, Robert Brown, as far back as 1827, the so-called Brownian motion. When tiny grains of a substance such as pollen are suspended in water and observed through a microscope, they are found to be in constant motion.

The peculiar thing about this motion is that it is quite random — the particles move in tiny rapid jerks, the direction of each little movement being entirely unrelated to the previous path of the particle. It had been suggested that this movement was really due to the incessant bombardment of the particle by water molecules, and Perrin took this suggestion one step further. It seemed to him that this collection of small particles could themselves be regarded as behaving like very large molecules and that they should then move and distribute themselves like the molecules of one gas in thermal equilibrium with another. This idea was supported by the observation that the particles

did not all sink to the bottom of the container, but ultimately reached a 'sedimentation equilibrium' in which the density of particles increased in a regular manner from the top of the column of liquid to the bottom.

Now this type of behaviour is just what is predicted by the Kinetic Theory. Since the molecules of a gas have mass, one expects the action of gravity on a vertical column of gas to produce a higher density of molecules at the bottom than at the top, with a corresponding increase in pressure from top to bottom. Equilibrium will be reached when, for a thin 'slice' of gas at any height, the weight of the gas molecules in the 'slice' is equal to the difference between the pressure force on the underside of the 'slice' and the pressure force on the upper side where the pressure, due to the slightly lower density, is a little less. The Kinetic Theory enables us to express these ideas in mathematical form and thus to predict the exact manner in which the number of particles per unit volume increases down the column. The expression obtained in this way for the ratio of the numbers per unit volume at two particular heights turns out to depend on the mass of each particle, the difference in heights, the temperature of the gas and, significantly, on Avogadro's number.

By means of very careful measurements using suspensions of gamboge in which the particles were very uniform in size and mass, with diameters between 10^{-3} and 10^{-4} mm, Perrin found that the distribution of particles through the column of water was exactly as predicted by the Kinetic Theory, so that the proposition that they behaved just like large molecules in constant collision with the water molecules was entirely justified. At last there was direct and quantitative evidence of the action of molecules. Not only this, but Perrin's measurements made it possible to obtain the first reliable estimate of Avogadro's number, since all of the other quantities which affected the distribution of the particles could be measured. Lastly, Perrin was able to observe the average displacements per unit time of individual particles as they wandered through the liquid. Here also a comparison with the Kinetic Theory was possible, since the solution of this problem of the 'random walk' of a gas molecule had been achieved by Einstein and Smoluchowski. Once again Perrin's measure-

¹ These predictions can largely be arrived at using only the qualitative assumptions described above; for the most part it is not necessary to use any quantitative estimates of the actual dimensions, spacings and velocities of the molecules. It is of interest to note the orders of magnitudes of these. For example, for the nitrogen molecules which largely constitute air at room temperature and pressure, we now know that their diameters are about 2.4×10^{-10} m, their average speed is about 400 ms^{-1} , the average distance between them is 3.3×10^{-9} m and they travel an average distance of approximately 400 diameters between collisions, which occur some 4×10^9 times per second.