
# UNIT 09

## Atomic & Nuclear Physics

# CHAPTER 27

# Atomic Structure

## 27.1 INTRODUCTION

The existence of the atom is widely accepted but its incredibly small size is hard to comprehend. Only recently have scientists been able to see and photograph individual atoms. Every breath you take contains about $10^{24}$ atoms. The full stop at the end of this sentence is a million atoms wide.

People get very confused about atoms. They ask questions like these:

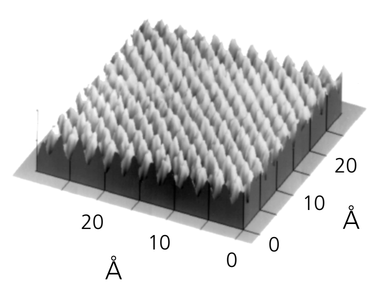
- If atoms are mostly empty space, how come a brick feels so hard?
- What colour is an atom?
- If the electron is negative why doesn't it get sucked into the positive nucleus?
- How many atoms are there in the universe? It must be a mind-bogglingly big number!
- If the nucleus is made up of positive particles, why don't they fly apart?
- How do we know atoms really exist if you can't see them?

Scientists have answered the last question but the rest need careful explanation. That's what this chapter is about.

## 27.2 FOUNDATIONS OF ATOMIC THEORY

The word *atom* comes from the Greek words *a* meaning 'not' and *tom* meaning 'to cut'; hence, not cuttable, or indivisible. This arose from the ideas of the Greek philosophers Democritus and Leucippus 2500 years ago.

**Photo 27.1**
Photo of surface of graphite in air taken by Dr K. Finlayson (Murdoch University, WA) using a scanning force microscope (SFM). Individual carbon atoms can clearly be seen at this magnification of 1.5 million times. The unit 'angstrom' (Å) is a non-SI unit equal to $10^{-10}$ m.




### Democritus

The word 'philosopher' was used differently from the way it is now. Until the word 'scientist' was coined in 1830, natural philosophers were people who loved learning about the world (it comes from the Greek *philos* = 'love', *sophia* = 'wisdom'). One of the first philosophers to suggest the idea of atoms was Leucippus; however, not much is known about his work. The earliest writing about atoms was that of **Democritus of Abdera** (460–371 BC). He argued that you could not keep cutting up something into smaller and smaller pieces forever; eventually you would end up with a piece that could not be cut any further — the 'atom' (Figure 27.1). Democritus also argued that atoms were in constant motion and that all atoms were composed of the same substance but differed in size and shape. His model accounted for many observable properties of matter but as he believed that the atom was the fundamental indivisible particle, he did not try to explain its structure.

Greek philosophers did not test their theories by experiments as scientists would today. This wasn't because they did not have the equipment to do so; Greeks had no inclination to conduct experiments because the philosophers came from élite (rich and powerful) families

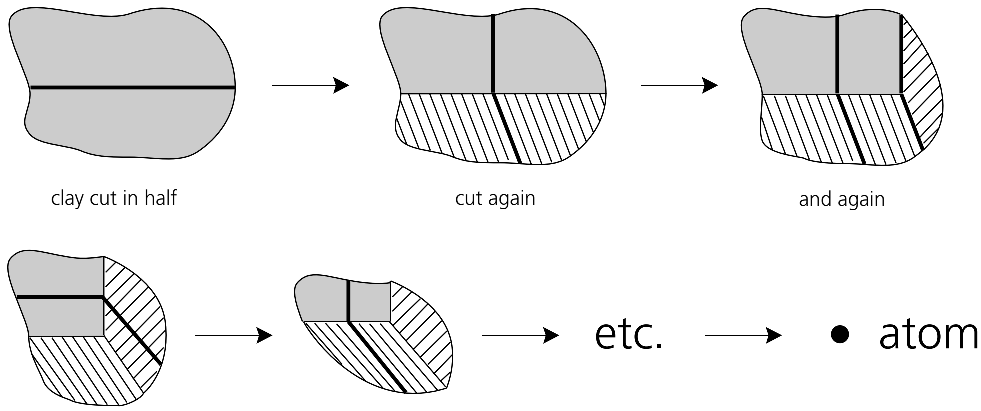



**Figure 27.1**
Democritus's concept of the atom.

**PHYSICS FACT**

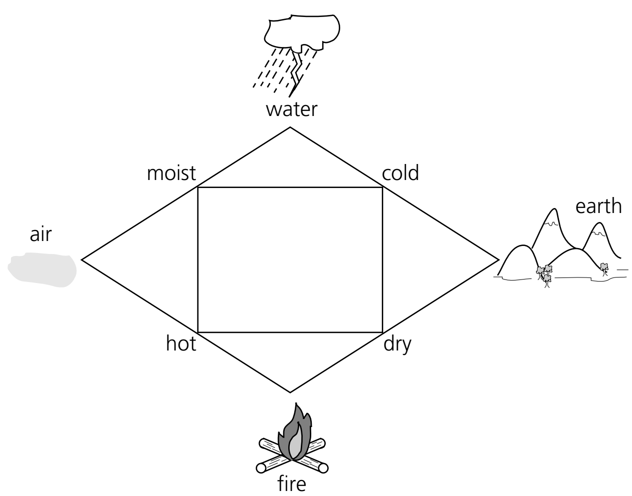
The word 'scientist' was coined by the distinguished Oxford professor William Whewell in 1834. Up until then they were called 'natural philosophers' (= lovers of learning about nature).

and thought manual work (such as experimenting) was only for slaves. They developed their ideas by reasoning and discussion. Their method of reasoning was to state some important principle or law, often based on observations of the heavens, then draw conclusions based on it. Experimentation generally did not occur — that was a seventeenth-century development.

## Aristotle

One of the most famous natural philosophers was **Aristotle**, born in 384 BC. His father was doctor to the king of Macedonia and Aristotle received a good education in Athens under the teaching of another famous Greek philosopher, Plato. His writings were vast and many of his theories went unnoticed until the thirteenth century, when Christian theologians began to endorse his work as being truth. Aristotle argued that matter could be divided an infinite number of times until there was a void, that is, nothing. He taught that matter was made up of four elements — earth, air, fire and water — and that different combinations produced different substances (Figure 27.2). This was at odds with the atomic theory but religious leaders could understand Aristotle's view of matter. They did not like the idea of atoms that moved around seemingly without the control of the gods. So from 300 AD to 1600 AD, the atomic theories lay dormant while Aristotle's ideas flourished.

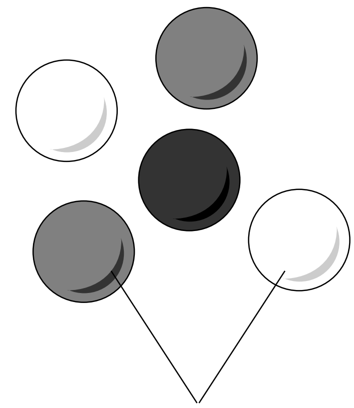
**Figure 27.2**
Aristotle's concept of the four elements.




# THE WORK OF JOHN DALTON 27.3

A long time passed before the idea of atoms was revived. In the seventeenth century, an Englishman by the name of Francis Bacon (1561–1626) introduced the idea of experimentation as a way of understanding nature. He reasoned that direct observation of nature rather than a study of Aristotle or theology (religious writings) gave a better idea of how the world worked. He is often thought of as the father of modern science. Later, people like Galileo and Descartes supported his idea of the experimental method. This inspired a lot of experimental work through England and Europe. Robert Boyle (1627–91) investigated the gas laws; Joseph Priestley (1733–1804) experimented with the extraction of gases; Antoine Lavoisier (1743–94) discovered the composition of air; and Henry Cavendish (1731–1810) discovered hydrogen. These experiments paved the way for a breakthrough in our understanding of matter. Near the beginning of the nineteenth century, the English scientist **John Dalton** (1766–1844) conducted a series of experiments, and published his atomic theory proposing the existence of individual particles called atoms in all matter (Figure 27.3), with a list of atomic masses. Dalton believed that all atoms of the same element were identical and that compounds were formed by the combination of atoms in small whole-number ratios. Other scientists went on to add to his theories and then in 1897, **J. J. (Joseph) Thomson** discovered the electron. This discovery led to the modern-day theory of atomic structure.

**Figure 27.3**
Dalton's model.




## 27.4 DISCOVERY OF THE ELECTRON

The big leap in our understanding of atomic structure came with the use of electricity in the laboratory. In the mid-nineteenth century, the effects of sparks, arcs and electrical discharges through gases were most interesting but of little importance. But after Heinrich Gessler invented the vacuum pump in 1855, electrical discharges through gases at low pressures produced brilliant results. Suddenly, the possibility of using vacuum tubes for electrical lighting (and making a fortune) was investigated and knowledge about the discharge increased dramatically. **Sir William Crookes** (1832–1919) in 1876 designed a number of tubes to study these charges. A variety of discharge tubes based on Crookes' designs are commonly available in physics classrooms today (Figure 27.4).

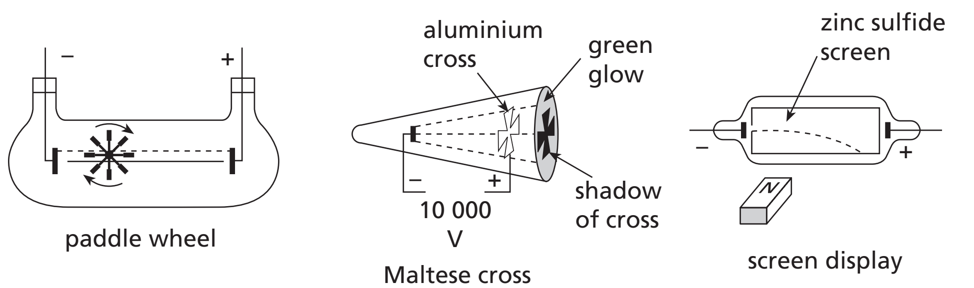
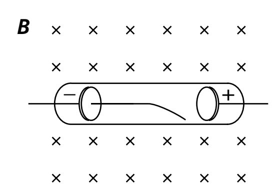



**Figure 27.4**
Discharge tubes commonly used in school laboratories.

These tubes contained various gases at low pressure and when a high voltage (about 20 000 V) was applied across the terminal at the ends, a purple light was seen; but as the pressure was reduced, the purple faded and the glass glowed with a green light near the positive end. There was dispute about what caused this glow but the invisible rays involved became known as 'cathode rays' as they emanated from the negative (cathode) terminal. The green light was an example of fluorescence (Latin *fluere* = 'to flow' and *esse* = 'to exist') — light given off by a substance (the glass) when being illuminated by energy from an external source (the discharge).

Crookes suggested that cathode rays would be deflected by magnetic fields (Figure 27.5) and by a series of experiments, Thomson was able to show this magnetic field deflection and so proved Crookes' hypothesis to be correct.

**Figure 27.5**
A magnetic field (into the page) deflects cathode rays downward.




Thomson devised a technique for passing cathode rays through an electric and a magnetic field that were orientated so as to exert opposing forces on the negatively charged rays. By this method, Thompson was able to measure the charge-to-mass ratio of the cathode ray particles, which he named electrons. The rays were deflected by the fields and struck the end of the glass tube, emitting light. The strength of the fields was then adjusted until the beam was not deflected. At this point the magnitude of the force exerted by the magnetic field, $\boldsymbol{F}_B$, was equal to the magnitude of the force exerted by the electric field, $\boldsymbol{F}_E$:

$$\boldsymbol{F}_E = \boldsymbol{F}_B$$
$$qE = \boldsymbol{B}q\boldsymbol{v}$$
$$\therefore\ \boldsymbol{v} = \frac{qE}{\boldsymbol{B}q} = \frac{E}{\boldsymbol{B}}$$

where $\boldsymbol{v}$ = the velocity of the electrons.

By knowing the strengths of the magnetic and electric fields, Thomson was able to calculate the velocity of the electron. When he switched off the electric field, he knew that the deflection of the electron was due just to the magnetic field. The curved path of the electron was due to centripetal force provided by the magnetic field.

The centripetal force formula:

$$\boldsymbol{F}_C = \frac{m\boldsymbol{v}^2}{r}$$

The magnetic force formula:

$$\boldsymbol{F}_B = \boldsymbol{B}q\boldsymbol{v}$$

(Recall these formulas from Chapters 5 and 25.)

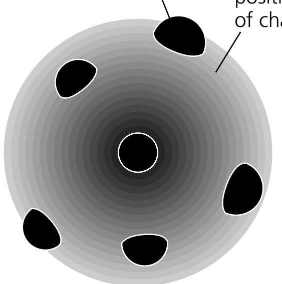
Hence:

$$\boldsymbol{B}q\boldsymbol{v} = \frac{m\boldsymbol{v}^2}{r}$$

or

$$\frac{q}{m} = \frac{\boldsymbol{v}}{\boldsymbol{B}r}$$

**Figure 27.6**
Thomson's 'plum pudding' model.




Substituting the value for velocity $\boldsymbol{v}$ above:

$$\frac{q}{m} = \frac{\frac{E}{\boldsymbol{B}}}{\boldsymbol{B}r} = \frac{E}{\boldsymbol{B}^2 r}$$

Thus the charge-to-mass ratio of the electrons would be given by:

$$\frac{q}{m} = \frac{E}{\boldsymbol{B}^2 r}$$

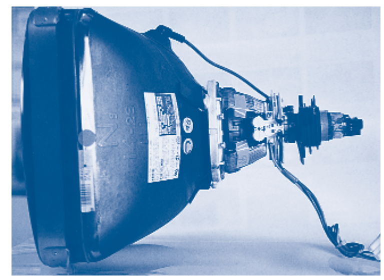
The ratio $q/m$ is more commonly referred to as the $e/m$ ratio, where $e$ stands for the charge on the electron. Thomson was able to measure the $e/m$ ratio to be $1.759 \times 10^{11}$ C kg$^{-1}$. However, neither $e$ nor $m$ can be determined individually by this method.

When Thomson determined the $e/m$ ratio for the hydrogen ion to be about 1800 times greater than that of the electron he realised that the electron was much smaller than the smallest known atom (hydrogen) and electrons were identical no matter where they came from. The only conclusion was that the electron was a sub-atomic particle; a conclusion later proved correct. Seven of Thomson's assistants went on to win Nobel prizes.

# 27.5 THOMSON'S MODEL OF THE ATOM

Thomson knew that an atom was electrically neutral but contained negatively charged electrons so he proposed a 'plum pudding' model of the atom (Figure 27.6). This envisaged the atom as a ball of positive charge (like a pudding) with electrons scattered throughout (like raisins). This model explained many features of the atom but couldn't explain others such as atomic spectra or radioactivity and was eventually replaced.

**Photo 27.2**
A cathode ray tube. The electron gun at the rear of a computer monitor can be seen.



## Modern cathode ray tubes

Cathode ray tubes (CRTs) have undergone continuous development since Crookes's original models. Many modern applications such as televisions, visual display units (VDUs) and cathode ray oscilloscopes (CROs) contain CRTs. In a modern CRT, a heated filament is used in an evacuated tube to produce an electron beam. The part of the tube that accelerates the high-speed electrons is called the **electron gun**.

## SR Activity 27.1 DISCHARGE TUBES

1. If you have an old TV set or computer monitor that is no longer any good, you may be able to take the cover off and look at the tube. Do not plug it in! And if it has been turned on in the past few months don't touch anything — the capacitors could still be charged (zap!!) Try to locate where the electrons are produced (the electron gun) and where the electric field and magnetic field coils could be. Share your findings with your class.
2. If your teacher demonstrates a flat screen discharge tube, suggest how you could test the effect of an electric field on it. Be careful not to stand too close to any operating discharge tube for too long. There is a slight danger from X-rays.

## Millikan's experiment

Thomson had worked out the ratio between charge and mass for the electron but it was not until the experiments of American physicist **Robert Millikan** (1868–1953) between 1909 and 1916 that the actual charge (in coulombs) and consequently the actual mass of the electron became known. Millikan's experiment was one of the classic experiments in physics.

A mist of oil was sprayed into the region above a pair of metal plates (Figure 27.7) and eventually a single oil drop fell through the hole in the top plate. When the plates were uncharged, the drops fell through at a steady velocity dependent on their weight. Being light objects, the oil drops reached a terminal velocity because they experienced considerable air resistance. By viewing the droplets through a microscope, Millikan was able to measure the terminal velocity and hence could determine the droplet's weight. When an electric field was applied between the plates, however, the motion of the drop was changed. It could be made to rise and fall depending on the voltage applied to the plates. A drop that remained stationary did so because it became charged during the spraying process and the electric force ($\boldsymbol{F}_E = \boldsymbol{B}qv$) was balanced by the gravitational force ($\boldsymbol{F}_W = m\boldsymbol{g}$). See Figure 27.8. If the upward electric force was greater than the gravitational force, then the charged droplet would move upwards. For a stationary drop:

**Figure 27.7**
Schematic diagram of Millikan's apparatus.

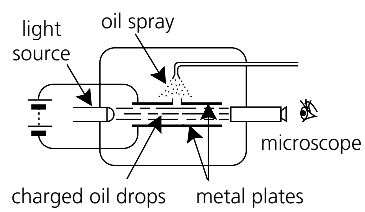
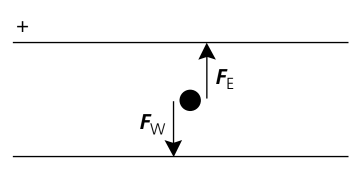
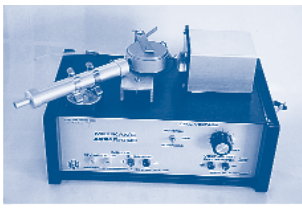



**Figure 27.8**
An oil drop balanced in the electric field between oppositely charged plates.




$$\boldsymbol{F}_E = \boldsymbol{F}_W$$
$$\boldsymbol{E}q = m\boldsymbol{g}$$
i.e.
$$q = \frac{m\boldsymbol{g}}{\boldsymbol{E}}$$

**Photo 27.3**
Millikan's apparatus — the telescope is on the left, the viewing chamber in the middle and the light source on the right.



Millikan studied the behaviour of thousands of oil drops and was able to work out the charges on them. He found that all charges were whole-number multiples of the minimum charge of $1.602\,1892 \times 10^{-19}$ coulombs. It is worth remembering that an uncharged oil drop still has billions of electrons but these are balanced by an equal number of protons. In Millikan's experiment, it is the *excess* electrons that are being counted.

Knowing the charge on the electron enables the mass of the electron to be calculated. If $e/m = 1.76 \times 10^{11}$ C kg$^{-1}$ and $e = 1.6 \times 10^{-19}$ C, then $m_e = 9.11 \times 10^{-31}$ kg.

## Example

In a Millikan oil drop experiment, an oil drop of mass $4.0 \times 10^{-16}$ kg was held stationary between a pair of electric plates held 2.0 cm apart. The voltage across the plates was 120 V. Assume $\boldsymbol{g} = 9.8$ m s$^{-2}$.

**(a)** Calculate the magnitude of the electric field between the plates when the oil drop was stationary.
**(b)** What was the size of the charge on the oil drop?
**(c)** How many elementary charges does this correspond to?

## Solution

**(a)**
$$E = \frac{\boldsymbol{V}}{d} = \frac{120}{0.02} = 6 \times 10^3 \text{ V m}^{-1}$$

**(b)**
$$q = \frac{m\boldsymbol{g}}{E} = \frac{4 \times 10^{-16} \times 9.8}{6 \times 10^3} = 6.5 \times 10^{-19} \text{ C}$$

**(c)**
$$\text{Number of elementary charges} = \frac{6.5 \times 10^{-19}\text{ C}}{1.6 \times 10^{-19}\text{ C}} = 4.$$

In 1910, Millikan was appointed Professor of Physics at the University of Chicago and in 1923 he was awarded a Nobel prize for this work.

At about the time that Thompson and Millikan were experimenting on electrons, the existence of radioactivity became known. The next section describes some of the milestones and people involved in its discovery.

## SR Activity 27.2 MATCHBOX MODELS

1. Take four matchboxes and put two marbles in one, three in the next, four in the next and five in the last. If you don't have marbles, any small objects such as 20 cent coins will do.
2. Ask a friend to measure the mass of each box and challenge her to tell you how many marbles or coins are in each box. Can she also tell you the mass of an empty matchbox?
3. How does this relate to Millikan's experiment?

## SR Activity 27.3 COMPUTER SIMULATIONS

If your school has Millikan's apparatus, your teacher may demonstrate how it works and collect some data. If not, there are some excellent computer simulations available.
If you don't have access to a commercial package there are several available on the Internet. If you can download one of these, you should find that you get a good feel for Millikan's experiment.

## Questions

1 Briefly compare and contrast the methods used by the Greeks and modern scientists in their investigation of scientific problems.

2 In a Millikan oil drop experiment, an oil drop of mass $1.05 \times 10^{-15}$ kg was held stationary between a pair of electric plates 2.6 cm apart and with a potential difference of 210 V. Assume $\boldsymbol{g} = 9.8$ m s$^{-2}$.

**(a)** Calculate the magnitude of the electric field between the plates when the oil drop was stationary.

**(b)** What was the size of the charge on the oil drop?

**(c)** How many excess electrons were on the drop?

# 27.6 THE DISCOVERY OF RADIOACTIVITY

Meanwhile, back to the debate about the nature of cathode rays — two opposing camps had developed. In England, Crookes led the particle theory group by saying the rays were torrents of negatively ionised gas molecules. Across the English Channel, the German scientist Heinrich Hertz led the opposition group, which said that as the rays could pass through metal foil they couldn't be big gas molecules and must be more like electromagnetic (light) waves. The breakthrough was soon to come by accident.

The history of science is littered with examples of scientists stumbling on amazing discoveries purely by chance. Does this mean that these serendipitous (lucky) discoveries are not real science? No! The discovery of radioactivity involved many of these fortuitous events and the talent of four main players — Wilhelm Roentgen, Henri Becquerel, and Marie and Pierre Curie.

## Wilhelm Roentgen

The breakthrough into the cathode ray wave versus particle debate came in November 1895. **Wilhelm Roentgen** (1845–1923), an obscure professor of physics at Würzburg, wanted to resolve this controversy so he bought a Crookes tube and soon found that something was happening outside the tube. Although his discharge tube was completely enclosed in black cardboard, he noticed that a piece of paper coated with the fluorescent compound barium platinocyanide, which happened to be lying on the bench, glowed while the tube was in operation. He identified the origin of the radiation in the glass wall of the tube where it was struck by the cathode rays. He didn't know what to call them so he just said they were **X-rays**. They were not charged particles since they were undeflected by magnetic fields. On 23 January 1896, in his one-and-only significant public lecture in Würzburg, Roentgen stated 'for the sake of brevity, I should like to use the term "rays" and to distinguish them from others I shall use the name "X-rays"'. The anatomist von Kolliker, who was present at this meeting, had his hand X-rayed (Photo 27.4) — the audience was astonished. Roentgen's breakthrough in the debate about whether cathode rays were particles or waves came as the consequence of this apparently accidental discovery. Within a few days it was news all over the world; it was the subject of music hall jokes and within a few days almost every physicist was repeating the experiments to admiring audiences.

**Photo 27.4**
Radiograph of the hand of anatomist Albert von Kolliker, made at the conclusion of Roentgen's lecture and demonstration at the Würzburg Physical-Medical Society on 23 January 1896.

### EI Activity 27.4 WHAT ROENTGEN SAW

See if you are able to try the following experiment (under teacher supervision). Use one of the Crookes gas discharge tubes from the physics laboratory. Place some luminescent mineral (e.g. uranium salt such as uranyl nitrate) about 1 m away from the tube in a darkened room and note the effect. Now place the uranium salt under an ultraviolet light (black light). How does it compare? Can you see how Roentgen was misled?

## Henri Becquerel

The early workers advanced many ingenious theories for the origin of the X-radiation. At first they thought it might be connected with the fluorescence of the glass walls of the tubes. **Henri Becquerel** (1852–1908), a professor of physics in Paris, began investigating a variety of fluorescent materials to see if they also emitted these X-rays. All trials with various minerals, metal sulfides, and other compounds known to fluoresce on exposure to visible light gave negative results. Then he remembered he had a sample of potassium uranyl sulfate from 15 years earlier, which he exposed to a bright light and placed on top of a photographic plate wrapped in two sheets of black paper. After several hours the plate had darkened. Cool!

Becquerel soon found out that this amazing behaviour had nothing to do with the fluorescence of the uranium salt. On 16 February 1896 he was going to confirm his result by repeating the experiment (as scientists do) but it was a rainy day so he left the uranium salt on top of the wrapped plate in a drawer, hopefully not near his lunch. A few days later he developed the plate and to his amazement a sharp, clear image was present. The penetrating radiation had come from the uranium itself, not from fluorescence under bright light. He had discovered a new source of radiation but it was not identical to X-rays. The radiation was emitted spontaneously by the uranium and was nothing to do with outside influences such as sunlight. The rays became known as 'Becquerel rays', but he did not do any more research on them, preferring instead to carry out research on X-rays. He discovered the phenomenon of naturally occurring radiation, partly by good scientific deduction and partly by accident. Becquerel rays were later to become known as 'alpha rays'.

### EI Activity 27.5 WATCH THE LUMINESCENCE

If you can get hold of an *old* watch or clock dial with a luminescent dial, look at it under a microscope or with a 10 power (or more) magnifying glass. Luminescent paint contains radioactive radium in the pigment. In a dark room, once your eyes are used to the dark, you should be able to see the watch dial clearly. Under the lens you should be able to see tiny light flashes that are given off whenever a particle is emitted by the unstable isotopes and strikes another chemical (zinc sulfide, ZnS) in the paint. During the manufacture of many of these dials in the United States in the 1920s, the women who worked on them used to make a fine point on their paintbrushes by wetting them between their lips. Those who did eventually ended up with tongue cancer, and fifteen of them died. Luminous dials these days contain H-3, Kr-85, Pm-147, Tl-204.

## Marie and Pierre Curie

**Marya (Marie) Sklodovska** was born in Warsaw (Poland) in 1867 and after high school did a degree in physics, topping the class. While enrolled in a second degree (in maths) in 1894 she met **Pierre Curie** (1859–1906) and they married the following year. On the advice of Becquerel, Pierre suggested Marie undertake a PhD studying the new 'Becquerel rays' as they were known. She began testing uranium compounds and measured their activity by using a device called an electrometer invented earlier by Pierre and his brother Jacques. As radiation passes through the air it can remove electrons from the gas molecules, creating positive ions, which will be attracted to a negatively charged electrometer and alter its charge.

She soon discovered that some uranium minerals were four or five times as active as they should be on the basis of their uranium content. She concluded that new substances must be present. She introduced the term 'radioactivity' (Latin *radius* = 'ray', a spoke of a wheel) to distinguish this form of emission from X-rays. Marie and Pierre began separating the uranium ore pitchblende into various fractions. The Curies bought a truckload of pitchblende and boiled it up in large vats, distilling the vapour. In July 1898 she discovered a new element, which she named 'polonium' (after Poland), and in December she isolated radium.

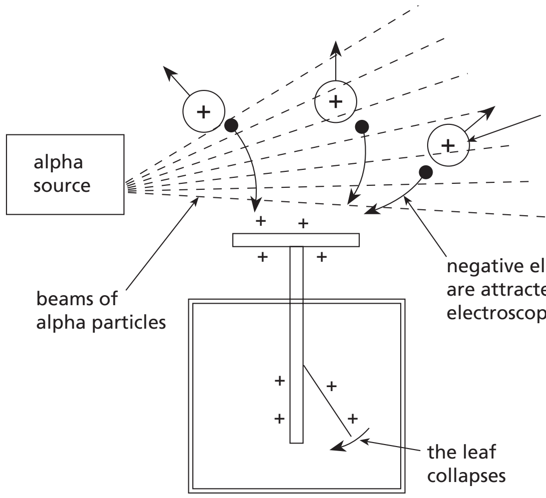



**Figure 27.9**
A radioactive source discharges an electrometer by ionising the air around it.

Unfortunately, Pierre was run over by a horse and dray and killed in Paris in 1905. Marie's daughter, Irene, joined the laboratory (The Radium Institute) and with her husband, Frederick Joliot, in 1934 was the first to demonstrate **artificial radioactivity** by bombarding aluminium with alpha rays. Later that year Marie died of leukaemia, a cancer of the blood, in her case caused by handling those radioactive substances for so long. When she died her skin was white from the lack of red blood cells. Today, in the garden at the rear of the Curies' laboratory, white roses grow — a moving metaphor and reminder of Marie's life.

Since then, many other radioactive elements have been discovered, although not all of them occur naturally. We now use the term 'nuclear radiation' instead of Becquerel rays because it is known that the radiation comes from the nucleus.

## EI Activity 27.6 RADIATION DETECTORS

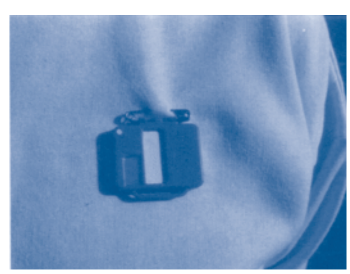
1 Charge an electroscope and time how long it takes to discharge into the air. What causes this? Repeat but hold a lighted taper close to the plate. Explain the result. Repeat again but, under teacher supervision, place an alpha particle source such as strontium-90 close to the plate. Explain this result.

2 Besides the electroscope, there are several other types of detectors available today, the simplest being photographic film that is often worn as a badge to monitor personal levels of radiation. Others include the Geiger-Müller tube, the thermo-luminescent dosimeter, the cloud chamber and the scintillation counter. Write a paragraph on one of the above types of detectors and include a drawing of what it looks like.

3 Why is a badge radiation monitor (see Photo 27.5) not the best way to protect you from radiation hazards?

4 Under teacher supervision, place a gamma source (such as cobalt-60) on top of a pinch of sodium chloride to irradiate it. After a while place the salt on top of a heated hotplate in a dark room and watch what happens. Explain how this is related to the thermo-luminescent dosimeter.

**Photo 27.5**
Radiation badge.



## Questions

3 Radioactivity was discovered by accident. Does this mean it is not 'real' science?

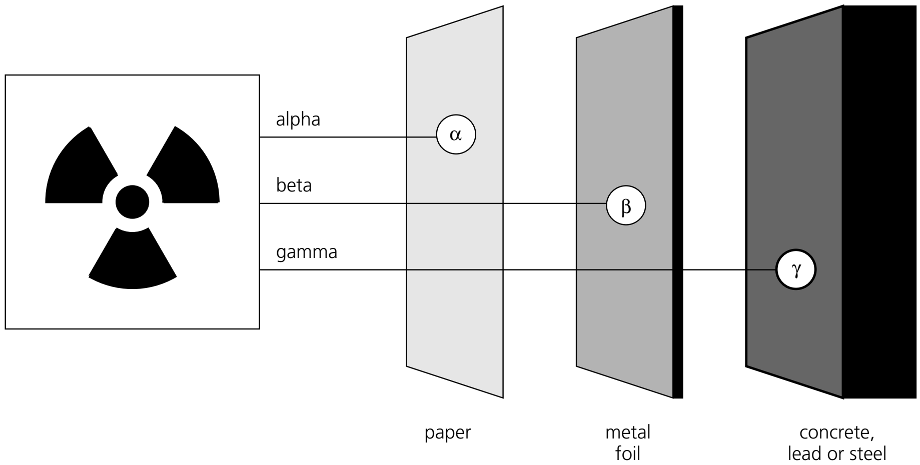
4 Radioactivity would never have been discovered, had it not been for the discovery of X-rays. Comment.

5 The measurement of nuclear radiation by the Curies depended on the invention of a particular instrument. Which instrument was this?

# NUCLEAR RADIATION 27.7

At the end of the nineteenth century research into radioactivity became one of the most important frontiers of research. One man led the field for many years. He was an experimental physicist, who won a Nobel Prize in chemistry and was perhaps the most productive physicist the world had known — the New Zealand born **Ernest Rutherford** (1871–1937). Rutherford won a scholarship to Cambridge to study under J. J. Thomson, beginning his life's work. He found that radiation could be classified according to its **penetrating ability** and in 1899 wrote: 'There are present at least two distinct types of radiation — one that is very easily absorbed, which will be termed for convenience $\alpha$ (alpha) radiation, and the other of a more penetrating radiation, which will be termed $\beta$ (beta) radiation' (after the first letters of the Greek alphabet). In 1900, the French physicist **Paul Villard** (1860–1934) discovered a third more penetrating radiation, which he called $\gamma$ (gamma).

**Figure 27.10**
The penetration power of radiation.

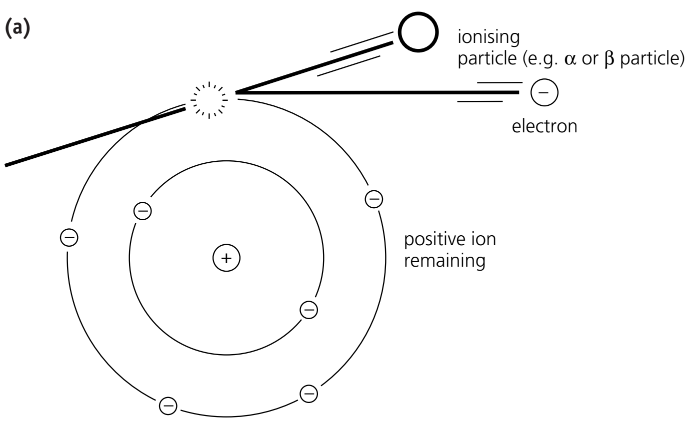


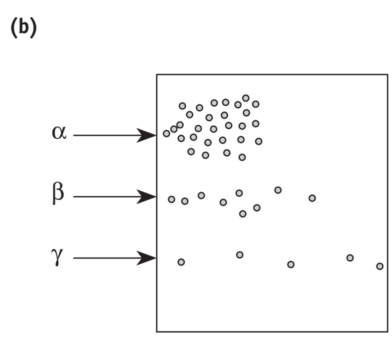

In 1903, Rutherford prepared some data on the relative penetrating power of the radiation. He used the term 'half-thickness' to indicate the thickness of aluminium that would reduce the intensity of the radiation to half.

- Alpha radiation is almost totally absorbed by a piece of paper; it will only travel a few centimetres in air. Aluminium half-thickness = 0.0005 cm.
- Beta radiation is almost totally absorbed by a millimetre of lead or a few millimetres of Perspex; it will travel about 1 m in air. Aluminium half-thickness = 0.05 cm.
- Gamma radiation is never fully absorbed and requires several centimetres of lead or over 1 m of concrete. Aluminium half-thickness = 8.0 cm.

## Ionising ability

**Figure 27.11**
**(a)** The process of ionisation.
**(b)** Comparative ionising ability.



Sometimes so much energy is absorbed by an atom that an electron completely escapes from the atom and a positive ion is produced. When this radiation causes ions to form, it is called **ionising radiation** (Greek *ion* = 'to go', hence the idea of ions moving around).

- Alpha radiation is powerfully ionising.
- Beta radiation has lower ionising ability than alpha radiation.
- Gamma radiation is only very weakly ionising.

The next step for physicists was to uncover the exact nature of each type of radiation. They did this by several ingenious experiments.

## Magnetic field effects

At the same time as radioactivity was discovered, the effect of magnetic fields on cathode rays was being investigated. It was a logical extension to try similar experiments on alpha, beta and gamma rays.

It was found that three types of deflections occurred when the radiation from a sample of radioactive material was passed through a magnetic field. In Figure 27.12, the rays are shown moving vertically up the page as they pass through a magnetic field into the page (shown by the crosses). Some rays deflect to the left, some to the right and some pass through unaffected. If you apply your hand or palm rules for the effect of a magnetic field on a charged moving particle, you will find that the rays passing to the left must be positively charged (alpha). The small curvature of the alpha rays indicates that they are relatively heavy particles. Similarly, rays to the right must be negatively charged (beta, or electrons), and rays passing straight through must be uncharged (gamma). In the early experiments, the evidence for deflection was collected on a photographic plate. In 1900, Rutherford and his assistant Thomas Royds showed that alpha particles are helium ions.

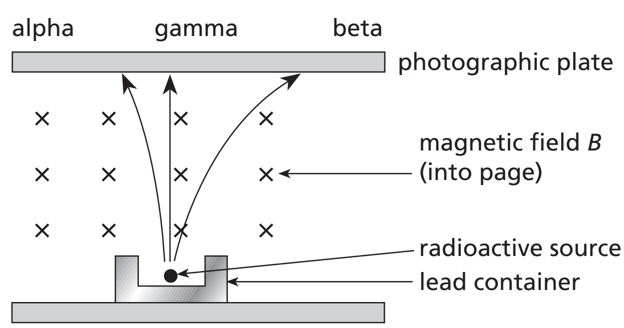



**Figure 27.12**
The deflection of alpha, beta and gamma rays by a magnetic field.

## Questions

**6** Alpha particles have a positive charge. How can they form neutral helium atoms in the tube? Where must the electrons come from?

**7** How did investigations of cathode rays lead to the discovery of radioactivity?

**8** Explain why alpha radiation has a low penetrating ability but is a powerfully ionising radiation, whereas gamma radiation is highly penetrating but a very weakly ionising radiation.

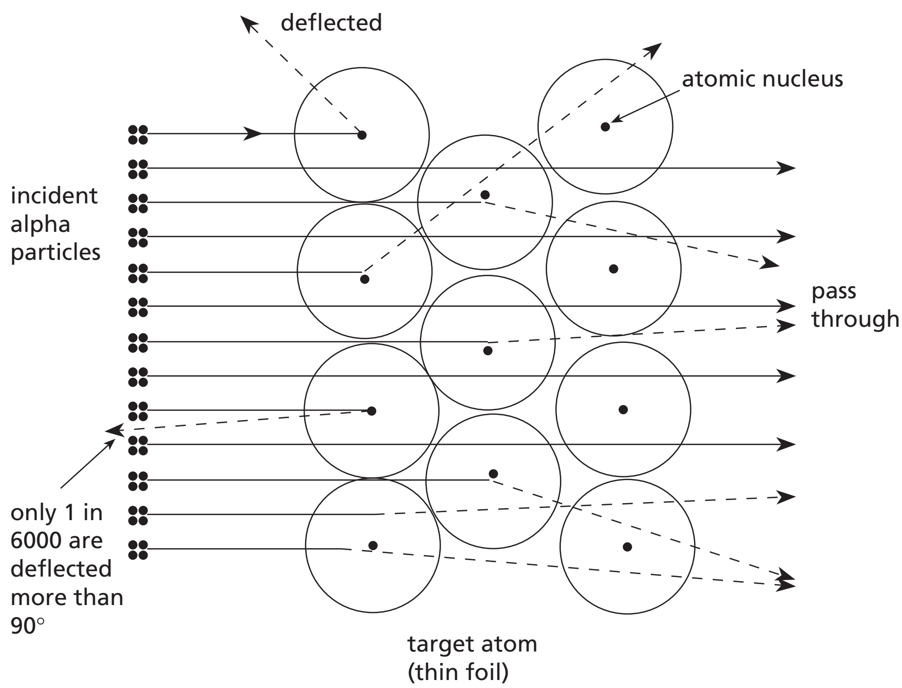
## The nuclear atom

The greatest discovery of all was yet to come. Rutherford and his colleague Hans Geiger had experimented on the scattering of alpha particles by thin pieces of heavy metal and found small deflections of about 1°. One day in 1910 at the University of Manchester, Geiger suggested to Rutherford that their young colleague Edward Marsden should begin research.

Rutherford jokingly said to let him see if any alpha particles scatter through a large angle, knowing that they probably would not. Rutherford told the story:

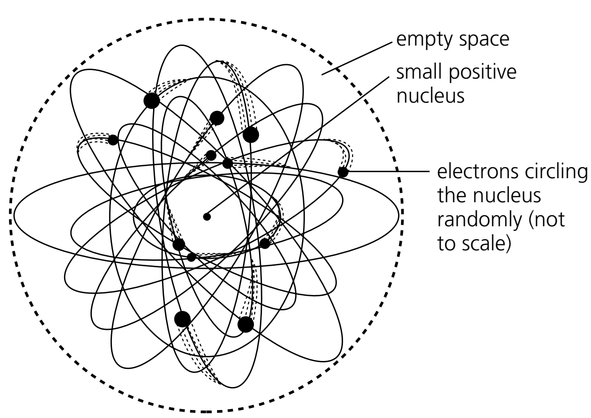
> Three days later Geiger came to me in great excitement saying, 'We have been able to get some of the alpha particles coming backwards'. It was the most incredible event that has ever happened to me in my life. It was almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you. (Source: E. Rutherford, *Philosophical Magazine*, Vol. 21, 1911.)

**Figure 27.13**
Alpha particles in a thin foil scattered by the atomic nucleus.




Geiger and Marsden used a beam of alpha particles from a radon source fired at gold foil a few atoms thick (about $10^{-7}$ m). Most of the alpha particles went straight through as though the metal was empty space and some scattered through small angles, but a few bounced straight back, deflected through 180°. Rutherford stated his nuclear model in 1911. The positive charge is concentrated in the massive centre of the atom with electrons revolving in orbits around it like planets around the sun (see Figure 27.14). Rutherford called this massive centre the **nucleus** (from an Indo-European language *knu* = 'nut, kernel'). He calculated that it had a diameter of about $10^{-14}$ m, about 1 ten-thousandth the diameter of the atom ($10^{-10}$ m).

**Figure 27.14**
Rutherford's model.



## Discovery of the proton

Once physicists had discovered alpha and beta particles, they developed a model for the structure of the nucleus. However, the nucleus of a hydrogen atom is too light to contain even a single alpha particle. In 1919, Rutherford bombarded some nitrogen atoms with $\alpha$ particles and succeeded in ejecting some positive particles. These particles are now known to be **protons** (Greek *protos* = 'first' or 'fundamental'), the nucleus of a hydrogen atom. These protons were soon found to have a positive charge, equal in magnitude but opposite in sign to the charge on an electron.

The work by Rutherford and other scientists provided evidence that protons were constituents of all atomic nuclei and he was able to propose a model for the nucleus as consisting of protons and much lighter electrons. Rutherford's laboratory is still radioactive to this day.

**PHYSICS FACT**

Famous artist Pablo Picasso said his paintings were inspired by Rutherford's theory that you could break up an atom into individual pieces and reconstruct it in different ways.

*Have you seen any of Picasso's portraits in which body parts have been swapped around?*

## Discovery of the neutron

In the following year, Rutherford delivered a lecture with far-reaching consequences. He proposed that it was possible that a proton in the nucleus might combine with an electron to form a neutral particle, which he called a **neutron** (Greek *ne* = 'not', *uter* = 'one or the other'; hence not positive or negative). This could be the particle to explain why the hydrogen nucleus has a charge of 1+ and a mass of 1 unit whereas helium has a charge of 2+ but a mass of 4 units. Over the next 12 years researchers looked in vain but their means of detecting particles was by deflection in magnetic and electric fields. As the neutron was predicted to be uncharged, it was unlikely to be detected.

In 1930, two German physicists, W. G. Bothe and Hans Becker, bombarded boron and beryllium with $\alpha$ particles and found a strange radiation — one that could penetrate lead better than gamma rays (and that's highly penetrating!). In France, Irene Curie (daughter of Marie and Pierre Curie) and her husband Frederick Joliot directed $\alpha$ particles onto a block of paraffin wax, a material rich in H atoms (see Figure 27.15). (You would recognise paraffin wax as surfboard wax — perhaps it was an overcast day and Irene didn't go surfing.) As the radiation struck the wax, 'knock-on' protons were formed on the other side. This seemed as though a new type of radiation was coming from the beryllium.

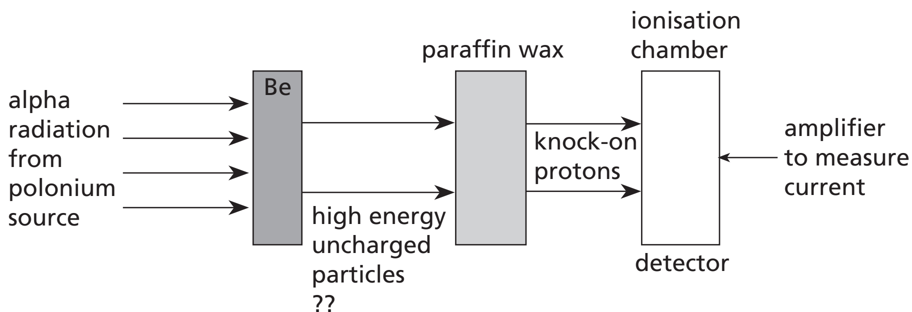



**Figure 27.15**
The apparatus used by Chadwick to identify the neutron.

In 1932 **James Chadwick**, while working at the Cavendish laboratory in Cambridge under Rutherford, used an ionisation chamber to measure the energies of the knock-on protons. Chadwick showed that the energy of the radiation was much higher than previously measured for $\gamma$ rays. The radiation could not be $\gamma$ rays but was instead due to uncharged particles of mass similar to that of the proton. These uncharged particles were neutrons. Chadwick was awarded the Nobel prize in 1935 for his discovery. During the war he headed Britain's team of nuclear physicists who worked on the atomic bomb. He was knighted in 1945.

## Questions

**9** Comment on the truth of this statement: 'In 1919, Rutherford proposed a model of the atom comprising a small positive nucleus of protons and neutrons with electrons circling it.'

# DIFFICULTIES WITH THE RUTHERFORD ATOM 27.8

The model of the atom as proposed by Rutherford provided a great step forward, but it also raised some difficulties. It is known that when a charged particle changes velocity (either speed or direction) it will radiate energy. If the electrons are in orbit around the positively charged nucleus then they should radiate energy because they are continually changing direction. If this were true, the atom would continuously lose energy and spiral into the nucleus. But it doesn't — no energy is radiated and it doesn't spiral inward. Something was wrong with Rutherford's model.

Some sort of modification was needed, and in 1913 it was the Danish physicist **Niels Bohr** who provided it by adding an essential idea — the quantum theory (from the Latin *quantus* = 'how much'). This raised a storm of controversy among scientists and philosophers. Bohr said that electrons could only exist in certain fixed orbits around the nucleus (Figure 27.16). He postulated that an electron in each orbit would move without losing energy or emitting radiation but when an electron jumped from one orbit to a lower one a single photon (a quantum) of light was emitted. Chapter 29 looks at the Bohr model and these spectral emissions in more detail.

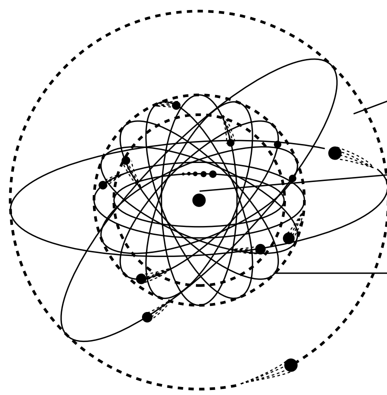



**Figure 27.16**
Bohr's model.

## Questions

**10** American physicist Douglas Giancoli remarked that students often say 'Don't tell us about the history and all the theories that were wrong. Tell us the facts as they are known today'. Critically evaluate this statement using examples from history.

**11** If the Thompson model had been correct, how would the results of the Rutherford scattering experiment have been different?

# THE MODERN STRUCTURE OF THE ATOM 27.9

We now know that an atom is made up of a dense nucleus composed of protons and neutrons and is surrounded by a cloud of electrons. About 99% of the mass is concentrated in the nucleus. If you had a nucleus the size of a strawberry it would bore its way through the ground. In Chapter 29 you will see that protons and neutrons have an internal structure — they are made up of particles called 'quarks'.

## Chemical characteristics

The number of protons in an atom determines its name. Hydrogen has one proton, helium has two, uranium has ninety-two and so on. The number of protons is called the atom's **atomic number**, which has the symbol *Z*. Particles residing in the nucleus (protons and neutrons) are called **nucleons**. The different types of nuclei are referred to as **nuclides**; for example, there is a hydrogen nuclide, a helium nuclide and so on.

## Mass and size

A proton or neutron has a mass 1836 times that of the electron, so in a hydrogen atom, for example, 99.95% of the mass is contained in the nucleus. However, the nucleus occupies only a tiny volume — about 1 trillionth the volume of the atom. If the nucleus were an orange seed, the outer region of the electron cloud would be 100 m away. How big is the electron? Well, the electron is considered to be a 'point' charge, so is not considered to have any size at all. Table 27.1 shows the masses and charges of some atomic particles.

### Table 27.1 THE MASS AND CHARGE OF THE ATOMIC PARTICLES

| PARTICLE | MASS (kg) | CHARGE (coulomb) |
|---|---|---|
| Proton | $1.6726 \times 10^{-27}$ | $+1.6 \times 10^{-19}$ |
| Neutron | $1.6750 \times 10^{-27}$ | 0 |
| Electron | $9.1095 \times 10^{-31}$ | $-1.6 \times 10^{-19}$ |

The simplest atom is that of hydrogen (Greek *hydro* = 'water', *genes* = 'to form'). It has one proton and no neutrons so its atomic number *Z* (number of protons) is 1 and its atomic mass *A* (p + n) is also 1. As it is an atom it is electrically neutral so it must have the same number of positive and negative charges; hence it has one electron. The number of protons determines the name of the atom; for example, all atoms with eight protons are called oxygen. If oxygen has eight electrons it is a neutral atom; if it has more electrons (e.g. 10) it is a negative ion ($O^{2-}$); fewer electrons (e.g. 6) makes it a positive ion ($O^{2+}$).

## Isotopes

As more radioactive elements were discovered between lead (atomic number 82) and uranium (92), it became increasingly difficult to see how they could all fit into the ten atomic number spaces. Cases began to appear in which two radioactive elements, distinctly different in the type of radiation they emitted, were otherwise chemically identical. For example, in 1913 Frederick Soddy found that there were two forms of lead; one was the ordinary sort with a mass of 207 units and the other, taken from a uranium deposit in Norway, had a mass of 206 units. Soddy concluded that such inseparable elements must occupy the same place in the periodic table and gave them the same **isotopes** (Greek *iso* = 'same', *topos* = 'place'). Isotopes of an element have the same number of protons but different numbers of neutrons. In the example above, both isotopes have 82 protons but one has 124 neutrons (206 – 82) and the other has 125 neutrons (207 – 82).

The chemical properties of isotopes are identical but the physical properties are different because of the different masses. For example, hydrogen has three isotopes called hydrogen (H), deuterium (D) and tritium (T) as shown in Figure 27.17. Their composition is shown in Table 27.2. Water made with ordinary hydrogen ($H_2O$) has a density of 1.000 g/mL and boils at 100.00°C, whereas 'heavy water' made from deuterium oxide ($D_2O$) has a density of 1.1079 g/mL and a boiling point of 101.42°C.

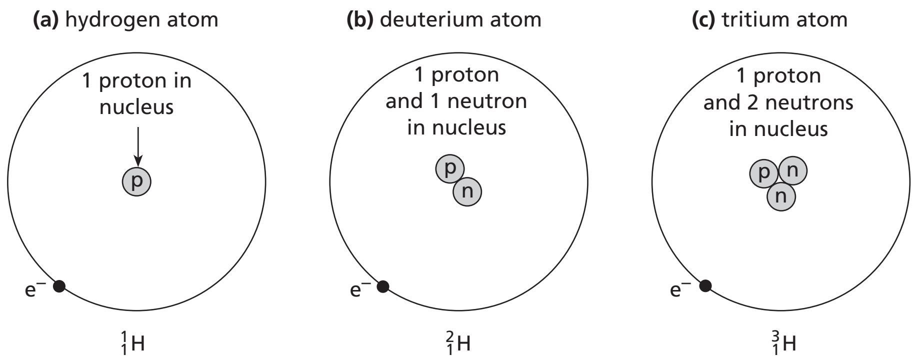



**Figure 27.17**
The isotopes of hydrogen: **(a)** hydrogen; **(b)** deuterium; **(c)** tritium.

## Table 27.2 THE ISOTOPES OF HYDROGEN

| ISOTOPE (NUCLIDE) | SYMBOL | PROTONS (ATOMIC NUMBER) | NEUTRONS | PROTONS + NEUTRONS (ATOMIC MASS) | ELECTRONS (= PROTONS) |
|---|---|---|---|---|---|
| Hydrogen | $^1_1H$ | 1 | 0 | 1 | 1 |
| Deuterium | $^2_1H$ | 1 | 1 | 2 | 1 |
| Tritium | $^3_1H$ | 1 | 2 | 3 | 1 |

Each isotope is referred to as a nuclide. The shorthand way of describing a nuclide is shown in Figure 27.18. For example, the isotopes of hydrogen are written as $^1_1H$, $^2_1H$ and $^3_1H$. The lead isotopes mentioned earlier would be written as $^{206}_{82}Pb$ and $^{207}_{82}Pb$ but can also be written as Pb-206 and Pb-207. (The atomic number 82 is not needed as all isotopes of lead have 82 protons.)

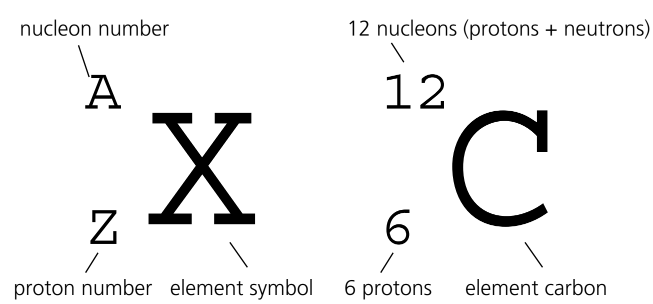



**Figure 27.18**
Notation of a nuclide. Can also be written as carbon-12 or just C-12.

Students often get confused when trying to work out symbols for nuclides. The most important thing to remember is that the atomic number $Z$ (the number of protons) and the symbol go hand-in-hand. All isotopes of lead have an atomic number of 82 and all atoms with an atomic number of 82 are lead. There can be other nuclides with an atomic mass of 206 or 207 besides lead, for example $^{206}_{84}Po$ or $^{207}_{83}Bi$. The symbol and the atomic mass can be found by consulting the periodic table shown in Appendix 6 or the table in Appendix 7. Note that the atomic mass of an element is the actual mass of the atoms of an element as it is found in nature. Since most elements exist as isotopes, this mass is really a weighted average of the masses of the isotopes present. For example, ordinary carbon consists of 98.89% carbon-12, 1.11% of carbon-13 and a trace ($1 \times 10^{-8}$%) of carbon-14. The average atomic mass is 98.89% of 12 plus 1.11% of 13 which equals 12.0111 units. This is called the **relative atomic mass** ($A_r$) of carbon. Only carbon-12 (the most abundant) has a relative atomic mass of exactly 12.0000.

### Example

Magnesium is found naturally occurring on Earth as three different isotopes. The abundance of each isotope is as follows: $^{24}_{12}Mg$, 78.70%; $^{25}_{12}Mg$, 10.13%; $^{26}_{12}Mg$, 11.17%. Calculate the relative atomic mass of magnesium.

### Solution

$$24 \times \frac{78.70}{100} = 18.888$$

$$25 \times \frac{10.13}{100} = 2.533$$

$$26 \times \frac{11.17}{100} = 2.904$$

Total (sum) = 24.32 units.

Most elements exist as a number of isotopes. Uranium contains three isotopes of atomic mass 234, 235 and 238. The particle composition of these isotopes is shown in Table 27.3. (The exact mass of a nuclide to four or more decimal places is in Appendix 8.)

**Table 27.3 THE ISOTOPES OF URANIUM**

| ISOTOPE | ATOMIC NUMBER | PROTON NUMBER | ELECTRON NUMBER | NEUTRON NUMBER | ATOMIC MASS | EXACT MASS (u) |
|---|---|---|---|---|---|---|
| | (Z) | (Z) | (Z) | (A – Z) | (A) | |
| U-234 | 92 | 92 | 92 | 142 | 234 | 234.04098 |
| U-235 | 92 | 92 | 92 | 143 | 235 | 235.04394 |
| U-238 | 92 | 92 | 92 | 146 | 238 | 238.05082 |

## Questions

You may need to refer to the periodic table to answer the following questions.

**12** Carbon has two common isotopes of atomic mass 12 and 14. Write their symbols.

**13** Helium exists in nature as two isotopes, of atomic mass 3 and 4. Prepare a table showing the particle composition of these isotopes.

**14** Neutrons are bigger than electrons but have higher penetrating power. Explain why.

**15** Write down the numbers of protons and neutrons in each of the nuclides represented by **(a)** $^{2}_{1}H$; **(b)** $^{12}_{6}C$; **(c)** $^{17}_{8}O$; **(d)** $^{23}_{11}Na$; **(e)** $^{32}_{16}S$; **(f)** $^{107}_{47}Ag$; **(g)** $^{127}_{53}I$; **(h)** $^{238}_{92}U$.

**16** What is the average atomic mass of silicon that consists of 92.22% Si-28, 4.70% Si-29 and 3.08% Si-30?

## 27.10 MASS DEFECT AND BINDING ENERGY

When accurate measurements of the masses of nucleons and nuclei are made, a significant discrepancy emerges. **The mass of a nucleus is always less than the combined individual masses of its constituent nucleons**. For example, the mass of a deuterium nucleus is $3.34364 \times 10^{-27}$ kg but the sum of the masses of the individual proton and neutron is $3.34760 \times 10^{-27}$ kg.

## Atomic mass units

The masses of neutrons, protons and electrons have been accurately measured using a mass spectrometer, in which the radius of curvature of a particle through known magnetic and electrostatic fields is determined (see Figure 25.30 in Chapter 25). The masses are expressed as 'unified **atomic mass units**', abbreviated *'amu'* or just *'u'*. **One unified atomic mass unit is one-twelfth the mass of an atom of the carbon isotope of atomic mass 12.0000 (that is, $^{12}$C). The unified mass unit includes the masses of the six electrons of the carbon atom.** In kilograms, 1 u = $1.6606 \times 10^{-27}$ kg.

The masses of atomic particles are as follows:

- proton ($m_p$) – 1.007 276 u
- neutron ($m_n$) – 1.008 665 u
- electron ($m_e$) – 0.000 549 u

A carbon-12 atom is made up of six protons, six neutrons and six electrons. The sum of these masses is as follows:

| | |
|---|---|
| 6 protons | 6.043 656 u |
| 6 neutrons | 6.051 990 u |
| 6 electrons | 0.003 294 u |
| Total mass | 12.098 940 u |

However, the mass of a $^{12}_{6}$C atom is 12.0000 u, which is less than the sum of its constituent particles. Where has the missing 0.098 940 u gone — into thin air?

The loss of mass was defined by scientist F. W. Aston as **mass defect** ($\Delta m$) and it represents the mass that has been converted to **binding energy**, energy that binds or holds the nuclear particles together. Holding six protons together in a nucleus is a difficult task as they are repelling each other. The binding energy is needed to overcome these huge electrostatic forces of repulsion. When we form a carbon nucleus we convert a mass of 0.098 94 u into binding energy. Einstein showed that mass and energy are equivalent and one can be converted into the other. Using his equation showing the relationship between mass and energy $E = mc^2$, we can calculate the amount of energy involved. The constant of proportionality is $c^2$ where $c$ is the speed of light in a vacuum ($3 \times 10^8$ m s$^{-1}$). The binding energy can be calculated:

$$E = 0.098\ 94 \text{ u} \times 1.6606 \times 10^{-27} \text{ kg} \times (3.00 \times 10^8)^2 = 1.48 \times 10^{-11} \text{ J}$$

Alternatively we could express the binding energy in units called **megaelectron volt** (MeV) where 1 u of mass defect is equivalent to 931 MeV. In the above case of carbon, a mass defect ($\Delta m$) of 0.098 94 u equals 0.098 94 u $\times$ 931 MeV/u = 92.1 MeV.

To break up the nucleus, you would have to expend the same amount of energy in the process.

The binding energy of a nucleus is the energy converted from mass when a nucleus is formed from its constituent nuclear particles, all initially in their free state.

### Example

For the nuclide $^4_2$He, calculate **(a)** the mass defect; **(b)** binding energy in MeV per nucleon.

### Solution

- Mass of $^4_2$He (nucleus and electrons) = 4.002 603 u.
- Components: 2 p = $2 \times 1.007\ 276$ u = 2.014 552 u
  2 n = $2 \times 1.008\ 665$ u = 2.017 330 u
  2 e = $2 \times 0.000\ 549$ u = 4.032 980 u.
- $\Delta m$ (mass defect) = 4.002 603 – 4.032 980 = 0.030 377 u.
- 1 unit of mass (1 u) = 931 MeV (megaelectronvolt) of energy.
  Hence: $^4_2$He binding energy = $0.030\ 377 \times 931$ MeV
  = 28.28 MeV

Obviously, the more protons present, the greater the repulsive forces and hence the greater the binding energy necessary. For example, C-12 has a binding energy of 92.1 MeV whereas U-235 requires 1781 MeV. However, expressing the required energy as the **binding energy per nucleon** gives a measure of how stable the nucleus is. The greater the binding energy per nucleon, the more stable it is.

For example, C-12 has 12 nucleons so 92.1 MeV ÷ 12 = 7.67 MeV/nucleon. U-235 has only 1781 ÷ 235 = 7.58 MeV per nucleon. The carbon nucleus is more stable.

## Variation in binding energy

If we plot a graph of the binding energy per nucleon against the atomic mass (see Figure 27.19), the most stable nuclides are apparent — those with atomic masses around 50–60 u, for example Fe-56. Those heavy nuclides with atomic masses to the right of this region are most likely to undergo nuclear fission and break into smaller nuclides to become more stable. Light nuclei, to the left of Fe-56, are more likely to undergo nuclear fusion and join together to form heavier nuclides. Fission and fusion are described in the next chapter.

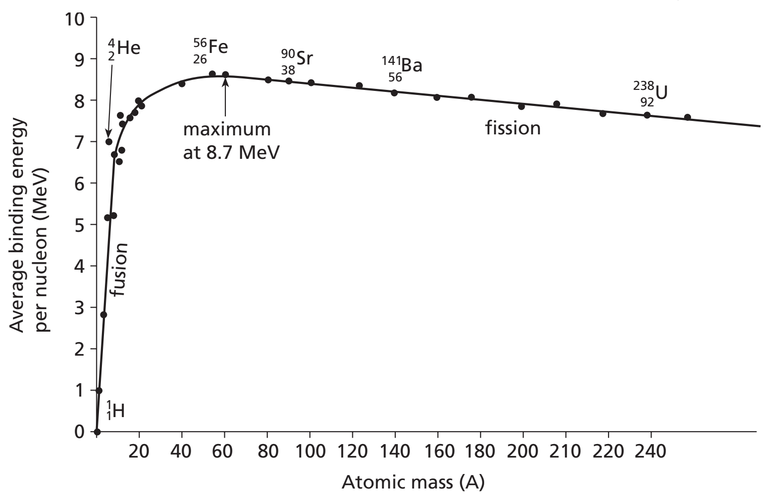



**Figure 27.19**
Graph showing the binding energy per nucleon as atomic mass changes.

## Question

**17** Use the table of exact atomic masses given in Appendix 8 to calculate the binding energy in **(i)** MeV and **(ii)** MeV per nucleon of **(a)** a tritium atom $^{3}_{1}H$, **(b)** helium-3 ($^{3}_{2}He$), **(c)** N-14, **(d)** O-16.

## Practice questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: * = low; ** = medium; *** = high.

### Review — applying principles and problem solving

***18** In a Millikan oil drop experiment, an oil drop of mass $1.077 \times 10^{-15}$ kg is held stationary between a pair of electric plates held 2.0 cm apart. The voltage across the plates is 110 V. Assume $\boldsymbol{g}$ = 9.8 m s$^{-2}$.
**(a)** What is the magnitude of the electric field between the plates?
**(b)** What is the charge on the oil drop? How many elementary charges is this equivalent to?

***19** Write down the number of protons and neutrons in **(a)** $^{207}_{82}Pb$; **(b)** $^{35}_{17}Cl$; **(c)** $^{15}_{7}N$; **(d)** At-215; **(e)** Bi-216.

***20** **(a)** Distinguish between *nucleon*, *nucleus*, *nuclide* and *neutron*. **(b)** Do *neutron* and *nucleus* come from the same Greek roots? If not, what is the difference?

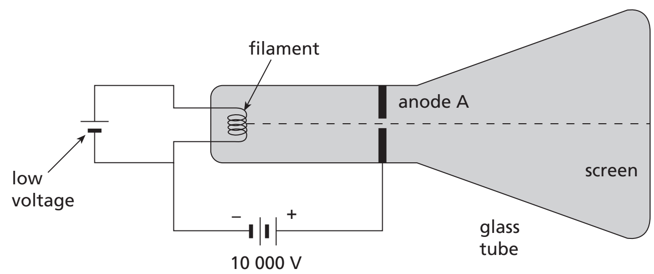
***21** Calcium has four isotopes with atomic masses of 40, 42, 43 and 45. **(a)** Write their symbols. **(b)** How many **(i)** neutrons and **(ii)** nucleons does each have?

**22 Calculate **(a)** the mass defect and **(b)** binding energy (MeV/nucleon) for **(i)** $^{35}_{16}S$; **(ii)** $^{235}_{92}U$; **(iii)** Cd-113; **(iv)** Li-7. See Appendix 8 for masses.

***23 Discuss the concept that progress in science depends on the development of the necessary equipment and technology to perform key experiments. Cite examples from history to illustrate your arguments.

**24 Figure 27.20 shows a simple arrangement of a cathode ray tube.
**(a)** Briefly explain why **(i)** the filament is heated; **(ii)** the electrons accelerate between the filament and the anode A.
**(b)** What effect, if any, is there on the number of electrons to the right of the anode if **(i)** the low voltage supply is increased; **(ii)** the high voltage supply is increased?

**Figure 27.20**
For question 25.



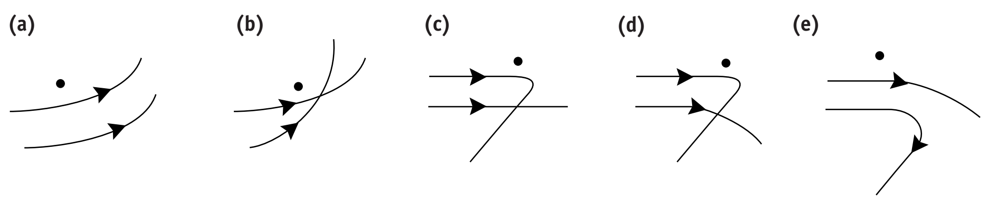

## Extension — complex, challenging and novel

***25 In a Millikan oil drop experiment, the potential difference between two large horizontal plates, 2.4 cm apart, is 1200 V. A tiny plastic sphere with a deficiency of three electrons is introduced into the electric field between the plates and it remains stationary. What is the mass of the sphere?

***26 Charged latex spheres can be used in Millikan experiments. One particular negatively charged sphere of mass $3.52 \times 10^{-14}$ kg is held stationary between two charged horizontal plates by an electric field of strength $9.8 \times 10^{4}$ V m$^{-1}$. How many excess electrons are there on the sphere?

**27 Figures 27.21**(a)** to **(e)** show paths of two $\alpha$ particles being deflected by a heavy nucleus. Which of the diagrams correctly represent(s) the deflection of two $\alpha$ particles of the same energy?

**Figure 27.21**
For question 28.




***28 An oil drop is held stationary between a pair of horizontal plates X and Y as shown in Figure 27.22**(a)**. The oil drop has a charge of $+6.4 \times 10^{-19}$ C. If the plates are turned through 90° as shown in Figure 27.22**(b)**, in which direction will the droplet now move?

**Figure 27.22**
For question 29.