Unit 6

Atomic Structure

This unit online, with its interactives:
science-playground.com/s/9a6

Nobody has ever seen inside an atom. An atom is so small that a single grain of salt holds several billion billion of them, and no light microscope will ever show one. Yet we know that an atom is almost entirely empty space, that its mass sits in a speck at its center tens of thousands of times smaller across than the atom, and that the electrons around it can only have certain energies, like a ladder whose rungs are fixed.

How do we know all this about something we can't see? The same way you can tell what is inside a closed box without opening it: by shaking it, tipping it, firing things at it and seeing what comes out. Over about 130 years, from John Dalton in 1803 to James Chadwick in 1932, scientists did exactly that to atoms. This unit follows their experiments, one picture of the atom at a time, and ends with the numbers that tell one atom from another: atomic number, mass number, atomic mass, isotopes, isobars and isotones.

6.1 Dalton's atomic theory

The idea of atoms is old. Around 400 BCE the Greek thinker Democritus argued that if you cut something in half again and again, you must at last reach a piece that can't be cut: an *atomos*, Greek for "uncuttable". In India, the Vaisheshika school of philosophy described tiny, eternal particles called **परमाणु** (paramanu). But these were ideas, not experiments. Nobody could test them.

That changed at the end of the 1700s, when chemists began to weigh carefully what went into a reaction and what came out. Two laws came out of their measurements:

- **The law of conservation of mass** (Antoine Lavoisier, 1789): in a chemical reaction, the total mass of the products equals the total mass of the reactants. Matter is neither made nor lost.
- **The law of constant proportions** (Joseph Proust, 1797): a compound always contains the same elements in the same proportion by mass. Pure water, from a Himalayan spring or a laboratory, is always 8 g of oxygen for every 1 g of hydrogen.

The English schoolteacher **John Dalton** (1766–1844) asked what matter must be like to obey these laws. He first presented his answer in 1803 and published it in full in *A New System of Chemical Philosophy* in 1808. It is called **Dalton's atomic theory**.

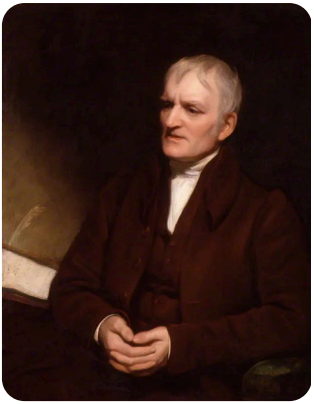


Figure 6.1 *John Dalton in 1835, painted by Thomas Phillips. Dalton taught mathematics and science in Manchester. He was also color-blind, and the first to describe color blindness scientifically; it was once called Daltonism.*
Photo: Thomas Phillips, Public domain

Postulates of Dalton's atomic theory

1. All matter is made up of tiny **indivisible** particles called **atoms**, and the atom is the smallest unit of an element.
2. Atoms of the **same element are alike** in mass and in their chemical and physical properties, and atoms of **different elements are different**.
3. Atoms combine with each other in **simple whole-number ratios** (1 : 1, 1 : 2, 2 : 3…) to form molecules.
4. In chemical reactions, atoms **keep their identity**: they are only rearranged into new combinations, and are **neither created nor destroyed**.

Each postulate explained something that had been measured. If atoms are never created or destroyed (postulate 4), and a reaction only rearranges them, the total mass can't change: that is conservation of mass. If every water molecule is the same group of atoms, each atom with its own fixed mass (postulates 2 and 3), every sample of water must have the same proportion by mass: that is constant proportions. Dalton also predicted a new law that he then checked himself, the **law of multiple propor-**

CAREFUL

Four postulates or seven?

CDC's chemistry book (Grade 11) gives Dalton's theory as these four postulates. Many notes split them into six or seven shorter points: atoms are indivisible; they can't be created or destroyed; atoms of one element are identical; atoms of different elements are different; they combine in simple whole-number ratios; an atom is the smallest particle that takes part in a chemical reaction. Each of those is part of one of the four above, so either list describes the same theory.

tions: when two elements form more than one compound, the masses of one that combine with a fixed mass of the other are in a simple whole-number ratio.

WORKED EXAMPLE

Multiple proportions in carbon's oxides

In carbon monoxide, 12 g of carbon is combined with 16 g of oxygen. In carbon dioxide, 12 g of carbon is combined with 32 g of oxygen. Show that these follow the law of multiple proportions.

Answer

For the same 12 g of carbon, the masses of oxygen are 16 g and 32 g.

$$16 : 32 = 1 : 2$$

A simple whole-number ratio, as the law says. Atoms explain why: carbon monoxide is one carbon atom with one oxygen atom (CO), and carbon dioxide one carbon atom with two (CO_2). There is no compound with one and a half oxygen atoms per carbon, because there is no half an oxygen atom.

Check 1

Which postulate of Dalton's atomic theory explains the law of conservation of mass?

a atoms of different elements have different masses

b in a reaction atoms are only rearranged, never created or destroyed

c atoms combine in simple whole-number ratios

d all matter is made of atoms

Modifications of Dalton's atomic theory

Dalton's theory was a huge step: for the first time, chemistry could be explained by particles with definite masses. But experiments over the next hundred years showed that some of its postulates are not quite right. Here is each postulate with what we know now.

Dalton's postulate	What we know now
1. Atoms are indivisible, the smallest unit of an element.	Atoms are made of smaller particles: **electrons**, **protons** and **neutrons** (6.2). But the atom is still the smallest particle that takes part in a chemical reaction: only its outer electrons are shared or passed on, and its nucleus is untouched.
2. Atoms of one element are alike in mass; atoms of different elements are different.	Atoms of one element can have different masses: **isotopes**, such as chlorine-35 and chlorine-37 (6.5). Atoms of different elements can have the same mass number: **isobars**, such as argon-40 and calcium-40. What really makes elements different is the **number of protons**.
3. Atoms combine in simple whole-number ratios.	Many compounds, especially carbon compounds, have large, far from simple ratios: table sugar is $C_{12}H_{22}O_{11}$, and a single starch molecule has thousands of atoms.
4. Atoms keep their identity, never created or destroyed.	True in **chemical** reactions. In **nuclear** reactions, atoms of one element change into atoms of another (**transmutation**), and a little mass turns into energy.

So the modified theory says: an atom is the smallest particle of an element that takes part in a chemical reaction; it is made of electrons, protons and neutrons; its identity is set by its number of protons; and atoms of one element may differ in mass.

> CAREFUL
> **"Modification" doesn't mean "wrong"**
>
> Dalton's theory isn't thrown away; it is still how every chemistry problem is done. Its picture of unchangeable, indivisible atoms holds perfectly for **chemical** reactions, which are all Dalton could study. The modifications come from physics experiments of energies far beyond any chemical reaction: electric discharges, radioactivity and nuclear reactions.

Check 2

The existence of isotopes shows that which postulate of Dalton's theory needs modifying?

a atoms are indivisible

b atoms combine in simple whole-number ratios

c all the atoms of one element are identical in mass

d atoms can be neither created nor destroyed

6.2 Subatomic particles

Introduction

The particles an atom is made of are called **subatomic particles**. There are three of them in every atom except ordinary hydrogen, which has no neutron: the **electron**, the **proton** and the **neutron**. Each was found by an experiment, and each experiment changed the picture of the atom.

The electron: cathode rays

In the 1870s, scientists found that a glass tube with most of its air pumped out glows when a high voltage is put across two metal plates sealed into its ends. The plate joined to the negative terminal is the **cathode**, the one joined to the positive terminal the **anode**. At a very low pressure (about 0.01 mmHg) and some thousands of volts, the glass at the far end from the cathode glows green, and an object placed in the tube casts a sharp shadow on the glow. Something was streaming from the cathode in straight lines. It was called **cathode rays**.

In 1897, at the Cavendish Laboratory in Cambridge, **J. J. Thomson** bent cathode rays with both an electric field and a magnetic field, and measured how much each bent them. The rays bent toward the positive plate, so they were **negatively charged particles**. From the bending he worked out the ratio of their charge to their mass, e/m: more than a thousand times larger than for the lightest ion known, hydrogen's (today's value is 1.76×10^{11} C/kg). And he got the same e/m whatever gas was in the tube and whatever metal the cathode was made of. So these particles were far lighter than any atom, and they were part of **every** atom. They were later called **electrons**.

Properties of cathode rays:

- They travel in straight lines from the cathode: an object in their path casts a sharp shadow.
- They are deflected toward the positive plate of an electric field, and by a magnetic field, so they carry **negative charge**.

- They make glass, and substances such as zinc sulphide, glow (**fluorescence**) where they strike.
- They can heat a thin metal foil they fall on, and pass through very thin foils.
- They produce **X-rays** where they hit a metal target (Wilhelm Röntgen, 1895).
- Their nature doesn't depend on the gas in the tube or the metal of the cathode.

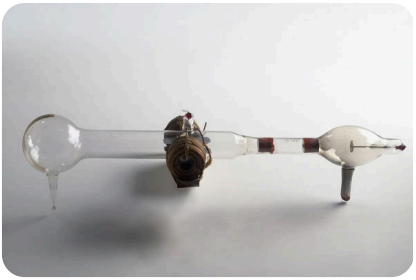


Figure 6.2 *The tube Thomson used in 1897 to discover the electron, now in the Science Museum in London. The cathode is in the small bulb on the right. The rays travel along the tube to the large bulb on the left, where they make the glass glow, and the coils wound round the tube make the magnetic field that bends them.*
Photo: Science Museum London / Science and Society Picture Library, CC BY-SA 2.0

The charge on one electron was measured by the American physicist Robert Millikan from 1909, by watching tiny charged oil drops float between two charged plates

CAREFUL

The paddle wheel

Some books say cathode rays turn a small paddle wheel placed in their path, which "shows they have mass and momentum". The wheel does turn, but later experiments showed it is pushed mainly by the heating of its vanes (the warmed side throws off gas molecules more strongly), not by the electrons' momentum, which is far too small. That cathode rays are particles with mass is shown by their bending in electric and magnetic fields. If an exam asks for properties, give the ones listed above.

(today's value is $e = 1.602 \times 10^{-19}$ C). Together with Thomson's e/m, that gave the electron's mass: today's value is 9.109×10^{-31} kg, about 1/1836 of a hydrogen atom's.

Since every atom is electrically neutral, an atom with negative electrons must also hold an equal amount of positive charge. Thomson pictured the atom in 1904 as a ball of positive charge with the electrons stuck in it, like the seeds in a watermelon (in England, the plums in a Christmas pudding, so it is called the **plum pudding model**). In 6.3 you will see how that picture was tested, and failed.

The proton: canal rays

In 1886 the German physicist **Eugen Goldstein** used a discharge tube whose cathode had holes (canals) in it, and found rays streaming backward through the holes, away from the anode. These **canal rays** (or anode rays) bent the opposite way to cathode

rays: they were **positively charged**. Unlike cathode rays, their e/m depended on the gas in the tube, because they are the atoms of that gas with electrons knocked off. The lightest, with the largest e/m, came from hydrogen: a hydrogen atom with its one electron removed.

In 1919 Ernest Rutherford fired alpha particles into nitrogen and found hydrogen nuclei coming out: they were inside the nitrogen nucleus too. In 1920 he proposed calling the hydrogen nucleus, the building block of positive charge in every nucleus, the **proton**.

The neutron

Protons alone couldn't account for the mass of atoms. A helium atom has 2 protons but about 4 times the mass of a hydrogen atom. Rutherford guessed in 1920 that the nucleus must also hold neutral particles. In 1932 **James Chadwick**, in Cambridge, fired alpha particles at beryllium. It gave off a radiation with no charge that knocked protons out of paraffin wax at high speed. From how hard it hit, he showed it was a stream of neutral particles with about the mass of a proton: the **neutron**. Chadwick was awarded the Nobel Prize in Physics in 1935.

Check 3

What did Thomson's finding that cathode rays have the same e/m in every gas show?

a that cathode rays are a kind of light
b that each gas has its own electrons
c that electrons are a part of every atom
d that electrons are positively charged

Properties of subatomic particles

Property	Electron	Proton	Neutron
Symbol	e^-	p^+	n^0
Discovered by	J. J. Thomson, 1897	E. Goldstein (canal rays), 1886; named by Rutherford, 1920	J. Chadwick, 1932
Where in the atom	around the nucleus	in the nucleus	in the nucleus
Charge (C)	-1.602×10^{-19}	$+1.602 \times 10^{-19}$	0
Relative charge	-1	$+1$	0
Mass (kg)	9.109×10^{-31}	1.6726×10^{-27}	1.6749×10^{-27}
Mass (u)	0.000549	1.00728	1.00866
Relative mass	1/1836 (taken as 0)	1	1 (a little more than a proton)

The unit **u** is the **unified atomic mass unit** (also written **amu**): one-twelfth of the mass of one carbon-12 atom, 1.6605×10^{-27} kg (6.5).

Three things to notice in the table:

- The proton's and electron's charges are **exactly equal and opposite**, so an atom with as many electrons as protons is neutral.
- The proton and neutron have **almost the same mass**, the neutron about 0.14% heavier.
- The electron is so light that **almost all the mass of an atom is in its nucleus**: 1,836 electrons would be needed to match one proton.

6.3 Rutherford's atomic model

The gold foil experiment

Ernest Rutherford (1871–1937), born in New Zealand, was working in Manchester on **alpha particles**: fast, positively charged particles shot out by radioactive elements such as radium. (An alpha particle is a helium nucleus, 2 protons and 2 neutrons, with a charge of +2; you will meet it again in Unit 12, Radioactivity.) In 1909 he asked

his assistant **Hans Geiger** and a 20-year-old student, **Ernest Marsden**, to fire alpha particles at a very thin metal foil and see whether any bounced back.

The setup:

- A radioactive source in a lead box with a narrow opening sent out a thin beam of alpha particles.
- The beam fell on a very thin foil of gold, less than a micrometer thick: Geiger gave one foil's thickness as 0.000086 mm, only about 300 atoms.
- Around the foil, a screen coated with zinc sulphide gave a tiny flash of light wherever an alpha particle struck it. The flashes were counted, in a darkened room, through a microscope that could be moved round the foil to any angle.
- Everything was inside an evacuated box, so that air didn't stop the alpha particles.

If Thomson's plum pudding model were right, the alpha particles should all pass through with hardly any deflection: the positive charge is spread thinly over the whole atom, so its push on a passing alpha particle is weak, and the light electrons can't turn a particle more than 7,000 times heavier than themselves.

What they saw:

1. Most of the alpha particles passed straight through the foil, undeflected.
2. Some were deflected through small angles.
3. A very few were deflected through large angles, and some bounced almost straight back. With a platinum plate, Geiger and Marsden found that about 1 in 8,000 of the alpha particles striking it bounced back off it.

Rutherford remembered it in a lecture in 1936: "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." Something in the foil was small, very heavy and strongly charged. He worked out the explanation and published it in 1911.

Rutherford's conclusions from the three observations:

Observation	Conclusion
Most alpha particles pass straight through.	Most of the atom is **empty space**.
A few are deflected through large angles.	The positive charge is not spread out but packed into a tiny region that repels the alpha particles strongly.
Very few bounce back.	That region, the **nucleus**, is extremely small (only a direct near-hit bounces back) and very heavy (an alpha particle bounces off it as a ball off a wall, not the other way round).

Postulates of Rutherford's atomic model

1. An atom has a very small, dense, positively charged center called the **nucleus**.
2. **All the positive charge and almost all the mass** of the atom are in the nucleus.
3. The nucleus is extremely small compared with the atom: its radius is about 10^{-15} m to 10^{-14} m, while the atom's is about 10^{-10} m, so the atom is tens of thousands of times wider than its nucleus.
4. The **electrons** revolve round the nucleus in circular paths called **orbits**, as planets go round the Sun. This is why it is also called the **planetary model**. The electrostatic attraction between the nucleus and an electron provides the force that keeps it in its orbit.
5. The number of electrons equals the number of units of positive charge on the nucleus, so the atom as a whole is **neutral**.
6. Most of the atom is empty space.

> DID YOU KNOW?
> **A grain of rice on a football ground**
>
> A gold atom is about 0.29 nm across, and its nucleus about 0.014 pm, roughly 20,000 times smaller. Blow the atom up to the size of a football ground, about 100 m across, and its nucleus would be a grain of rice on the center spot, holding more than 99.9% of the mass. Lighter atoms have even smaller nuclei for their size: a hydrogen atom is about 60,000 times wider than its proton.

Limitations of Rutherford's atomic model

1. **It can't explain why atoms are stable.** By the laws of electricity and magnetism, a charged particle moving in a circle is always accelerating and must give out energy as light. An orbiting electron should lose energy all the time and

spiral into the nucleus, in about a hundred-billionth of a second (for hydrogen, 1.6×10^{-11} s). Every atom would collapse. But atoms are stable: they last for billions of years.

2. **It can't explain the line spectra of atoms.** A spiraling electron would give out light of every color, smeared out. But heated hydrogen gives out light of only a few sharp colors (its line spectrum).
3. It says nothing about how the electrons are arranged round the nucleus.

Check 4

Which observation of the gold foil experiment shows that most of an atom is empty space?

- **a** most alpha particles passed straight through the foil
- **b** some alpha particles were deflected through small angles
- **c** a very few alpha particles bounced back
- **d** the zinc sulphide screen glowed

Check 5

Why does an orbiting electron in Rutherford's model fall into the nucleus?

- **a** it is attracted by the neutrons
- **b** moving in a circle, it gives out energy as light and loses energy
- **c** alpha particles knock it inward
- **d** it has no charge

A STORY FROM SCIENCE

From New Zealand to the nucleus

Ernest Rutherford was the fourth of twelve children of a farming and flax-milling family in New Zealand. In 1895 he won a scholarship that took him to the Cavendish Laboratory in Cambridge, where J. J. Thomson was about to discover the electron. Rutherford went on to explain radioactivity, and in 1908 he was awarded the Nobel Prize, in **Chemistry**, though his work was physics. His most famous discovery, the nucleus, came three years after the prize, and in 1919 he made the first artificial transmutation of one element into another, turning nitrogen into oxygen. He later led the Cavendish Laboratory, where Chadwick found the neutron.

6.4 Bohr's atomic model

In 1913 the Danish physicist **Niels Bohr** (1885–1962), who had worked with Rutherford in Manchester, kept Rutherford's nucleus but gave the electrons new rules. He took an idea from the newest physics of the time: that energy comes in fixed packets, **quanta**, as Max Planck and Albert Einstein had shown for light.

Postulates of Bohr's atomic model

1. Electrons revolve round the nucleus only in certain fixed circular orbits, called **stationary orbits**, **energy levels** or **shells**. They are numbered $n = 1, 2, 3, 4\ldots$ outward from the nucleus, or named K, L, M, N….
2. While an electron stays in one of these orbits, it **does not give out energy**. (This is the rule that saves the atom from collapsing.)
3. Each orbit has a **fixed energy**. The energy is lowest in the innermost orbit, K, and increases outward: $E_1 < E_2 < E_3 < \ldots$
4. Only those orbits are allowed in which the electron's angular momentum is a whole-number multiple of $\frac{h}{2\pi}$: $mvr = \frac{nh}{2\pi}$, where $n = 1, 2, 3\ldots$ is the orbit's number (the **principal quantum number**).
5. An electron gives out or takes in energy only when it **jumps from one orbit to another**: $\Delta E = E_2 - E_1 = h\nu$.



Figure 6.3 *Niels Bohr. He received the Nobel Prize in Physics in 1922 for his work on the structure of atoms.*
Photo: AB Lagrelius & Westphal, Public domain

In postulate 4, m is the electron's mass, v its speed, r the orbit's radius and h Planck's constant (6.626×10^{-34} J s). In postulate 5, an electron jumping inward, from a higher energy E_2 to a lower E_1, gives out the difference as one packet of light of frequency ν; to jump outward, it must take in exactly that much. A bigger jump gives more energy and a higher frequency: violet instead of red.

The energy levels of hydrogen

For the hydrogen atom, Bohr's rules give each orbit a radius and an energy. The radius grows as n^2: the second orbit is 4 times as wide as the first (0.053 nm), the third 9 times. The energy is

$$E_n = -\frac{13.6}{n^2}\ \text{eV}$$

where 1 **electronvolt** (eV) = 1.602×10^{-19} J is the energy an electron gains crossing 1 volt. The energy is negative because the electron is bound: it would take energy to pull it away to zero, outside the atom.

Bohr's orbits also gave the rule you used in Grade 8 and 9 for filling shells: each shell holds at most **$2n^2$** electrons (K 2, L 8, M 18, N 32), and electrons fill the lower-energy shells first. This arrangement of electrons in shells is sometimes called the **Bohr–Bury scheme**.

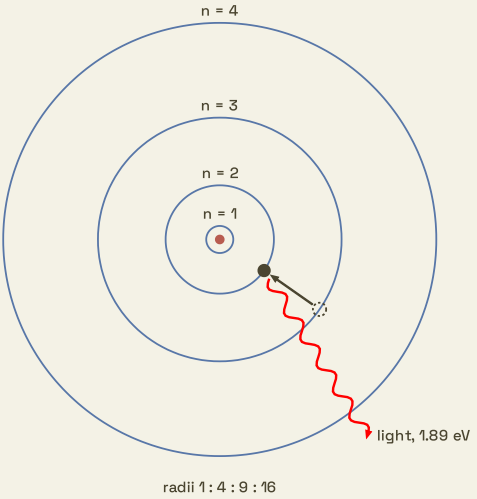
Merits and limitations of Bohr's model

Merits:

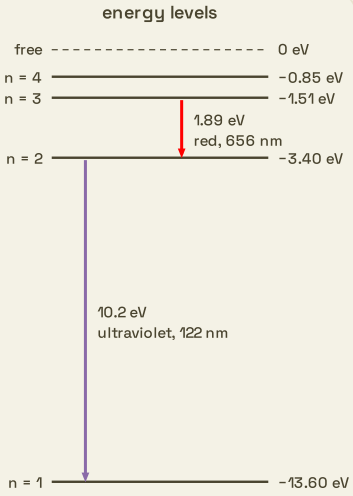
- It explains why atoms are stable: an electron in a stationary orbit doesn't radiate.
- It explains the **line spectrum of hydrogen**, and calculates the wavelengths of its lines.
- It calculates the size of the hydrogen atom and the energy needed to remove its electron (13.6 eV), both matching experiments.
- It gives the idea of electron shells, from which the $2n^2$ rule and electronic configurations came.

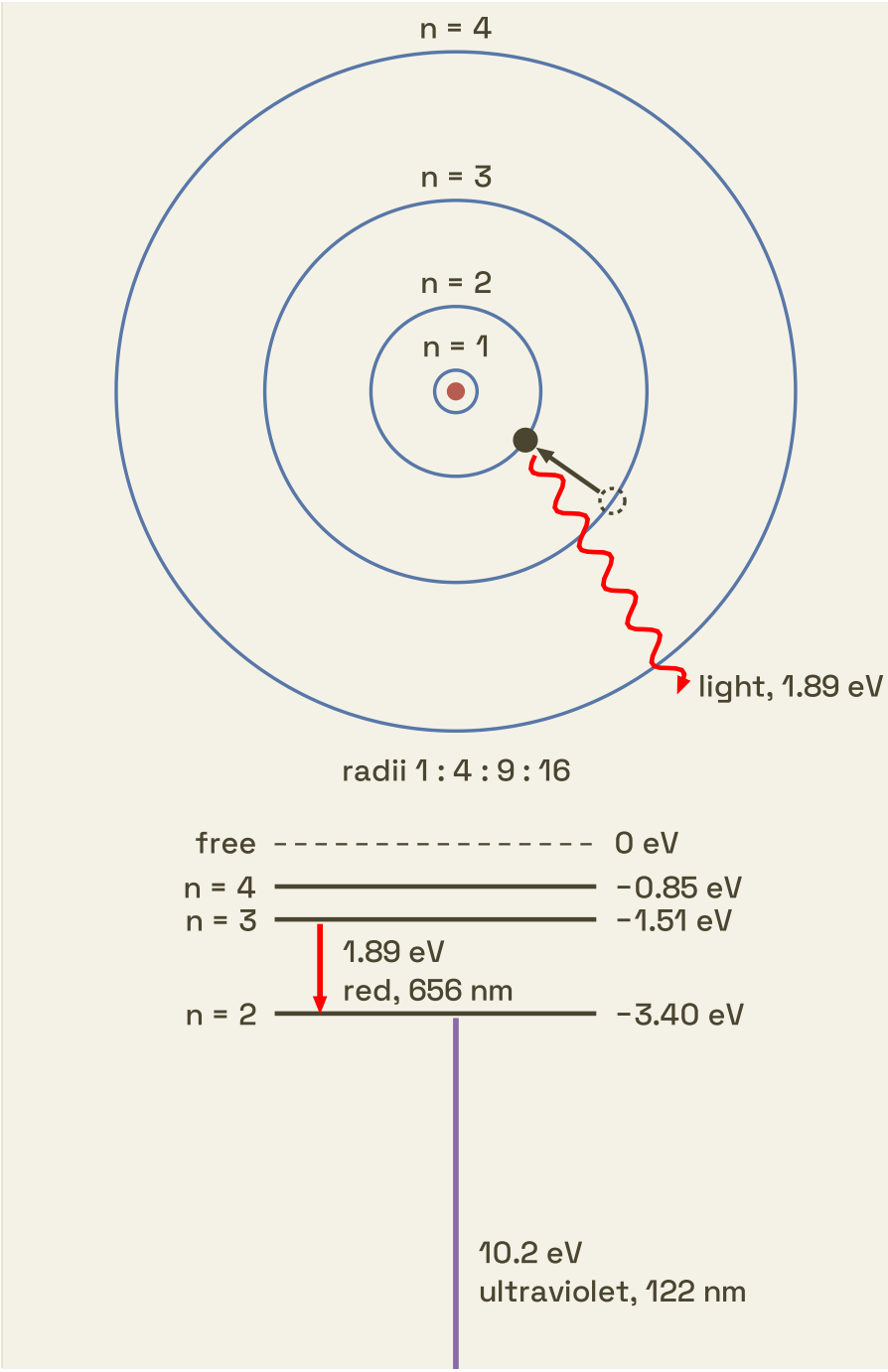
Limitations:

- It works only for hydrogen and other atoms or ions with **one electron** (He^+, Li^{2+}). For atoms with two or more electrons, its numbers come out wrong.
- It can't explain why some spectral lines, looked at closely, are pairs of lines, or why lines split in a magnetic field.
- It pictures the electron as a tiny ball on a fixed circular track. Since 1926 we know an electron has no definite path: we can only say where it is likely to be, an **electron cloud** (an **orbital**).


n = 4
n = 3
n = 2
n = 1
light, 1.89 eV

radii 1 : 4 : 9 : 16


energy levels
free
0 eV
n = 4
−0.85 eV
n = 3
−1.51 eV
1.89 eV
red, 656 nm
n = 2
−3.40 eV
10.2 eV
ultraviolet, 122 nm
n = 1
−13.60 eV


n = 4
n = 3
n = 2
n = 1
light, 1.89 eV
radii 1 : 4 : 9 : 16
free
0 eV
n = 4
−0.85 eV
n = 3
−1.51 eV
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red, 656 nm
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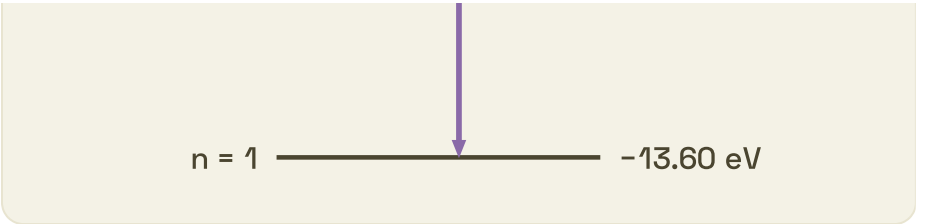


Figure 6.4 *The hydrogen atom by Bohr's model, both sides to scale. Left: the first four orbits, their radii in the ratio 1 : 4 : 9 : 16. Right: the same four energy levels, crowding together toward 0 eV. An electron dropping from the third level to the second gives out 1.89 eV, hydrogen's red light.*

WORKED EXAMPLE

Hydrogen's red light

How much energy does a hydrogen atom give out when its electron drops from the third orbit to the second?

Answer

$E_3 = -\frac{13.6}{3^2} = -1.51$ eV and $E_2 = -\frac{13.6}{2^2} = -3.40$ eV.

$$\Delta E = E_3 - E_2 = -1.51 - (-3.40) = 1.89 \text{ eV}$$

1.89 eV, which is $1.89 \times 1.602 \times 10^{-19} = 3.03 \times 10^{-19}$ J. Light with this much energy in each packet is red, of wavelength 656 nm, and that is exactly the bright red line in hydrogen's spectrum. Bohr's model was the first to calculate it.

Check 6

According to Bohr, when does an electron give out energy?

a all the time, as it moves round its orbit
b only when it is in the K shell
c when it jumps from a higher orbit to a lower one
d when it jumps from a lower orbit to a higher one

CAREFUL

Shells are energies, not race tracks

A drawing of an atom with neat circles is a way of counting electrons by energy level, and it is the picture your exams expect for configurations. But no electron goes round a track. The modern picture keeps Bohr's energy levels and his numbers n, and replaces his orbits with clouds of probability.

6.5 Concept of atomic number, atomic mass, isotopes,

isobars and isotones

Atomic number and mass number

The **atomic number** (Z) of an element is the number of protons in the nucleus of its atom. In a neutral atom it is also the number of electrons.

The atomic number is what makes an element that element. Every atom with 6 protons is carbon; give it a seventh and it is nitrogen. Elements in the modern periodic table are arranged in order of atomic number.

The **mass number** (A) of an atom is the total number of protons and neutrons in its nucleus. Protons and neutrons together are called **nucleons**.

$$A = Z + N \quad \text{so} \quad N = A - Z$$

where N is the number of neutrons. An atom is written with its mass number at the top left of its symbol and its atomic number at the bottom left:

$${}^{A}_{Z}\mathrm{X} \qquad \text{for example} \qquad {}^{23}_{11}\mathrm{Na},\ {}^{35}_{17}\mathrm{Cl}$$

or simply by name and mass number: sodium-23, chlorine-35, Na-23.

WORKED EXAMPLE

Counting the particles

Find the numbers of protons, electrons and neutrons in (i) ${}^{23}_{11}\mathrm{Na}$, (ii) ${}^{238}_{92}\mathrm{U}$ and (iii) the sodium ion, Na^{+}.

Answer

i. Protons = Z = 11; electrons = 11 (a neutral atom); neutrons = A − Z = 23 − 11 = **12**.

ii. Protons = 92; electrons = 92; neutrons = 238 − 92 = **146**.

iii. Na^{+} is a sodium atom that has lost one electron. The nucleus doesn't change: protons 11, neutrons 12, but electrons 11 − 1 = **10**.

Atomic mass

Atoms are far too light to weigh in grams: a carbon atom has a mass of about 2×10^{-23} g. So chemists compare their masses with a standard atom instead. Since 1961 the standard has been the **carbon-12 atom**, given a mass of exactly 12 u.

The **atomic mass unit** (amu, or u) is one-twelfth of the mass of one carbon-12 atom: $1\ \mathrm{u} = 1.6605 \times 10^{-27}$ kg.

YOUR TURN

How many protons, neutrons and electrons are there in (i) $^{40}_{20}\mathrm{Ca}$, (ii) $^{31}_{15}\mathrm{P}$ and (iii) the chloride ion, $^{35}_{17}\mathrm{Cl}^-$?

Answer

i. 20 protons, 20 neutrons, 20 electrons.

ii. 15 protons, 16 neutrons, 15 electrons.

iii. 17 protons, 18 neutrons, 18 electrons (the ion has gained one electron).

A STORY FROM SCIENCE

Moseley's X-rays

At first, "atomic number" meant only an element's place in the periodic table. In 1913 the young English physicist Henry Moseley fired electrons at targets of many elements and measured the X-rays they gave out. The X-rays' frequency rose in perfect steps from one element to the next, and he showed that each step was one more unit of positive charge on the nucleus: the atomic number is the number of protons. His measurements showed where the table still had missing elements, at 43, 61, 72 and 75, all found later. Moseley was killed in the First World War at Gallipoli in 1915, aged 27.

The **relative atomic mass** (or **atomic mass**, or atomic weight) of an element is the average mass of its atoms compared with one-twelfth of the mass of a carbon-12 atom:

$$\text{atomic mass} = \frac{\text{average mass of one atom of the element}}{\frac{1}{12} \times \text{mass of one carbon-12 atom}}$$

As a ratio of two masses it has no unit, but it is the same number as the atom's mass in u. Oxygen's atomic mass is 16.00: an oxygen atom is, on average, 16 times as heavy as one-twelfth of a carbon-12 atom.

Mass number and atomic mass are not the same thing. The mass number counts particles, so it is always a whole number. The atomic mass is a measured mass, and it is almost never a whole number: the mass of a chlorine-35 atom is 34.969 u, not 35, because protons and neutrons don't weigh exactly 1 u each, and a little of their mass goes into binding the nucleus. And an element's atomic mass is an **average** over its isotopes, which brings us to them.

Isotopes

Isotopes are atoms of the same element that have the same atomic number but different mass numbers. They have the same number of protons and electrons but **different numbers of neutrons**.

Hydrogen has three isotopes, the only ones with names of their own:

Isotope	Symbol	Protons	Neutrons	Electrons
protium (ordinary hydrogen)	${}^{1}_{1}\mathrm{H}$	1	0	1
deuterium (heavy hydrogen)	${}^{2}_{1}\mathrm{H}$ or D	1	1	1
tritium	${}^{3}_{1}\mathrm{H}$ or T	1	2	1

More examples: carbon-12, carbon-13 and carbon-14 (6 protons with 6, 7 and 8 neutrons); chlorine-35 and chlorine-37 (17 protons with 18 and 20 neutrons); uranium-235 and uranium-238 (92 protons with 143 and 146 neutrons).

Properties of isotopes:

- They have the **same chemical properties**, because chemistry depends on the electrons, and isotopes have the same number of electrons with the same arrangement. Water made with deuterium, D_2O (heavy water), still puts out fires and dissolves salt.
- They have slightly **different physical properties** that depend on mass: density, melting and boiling points, and how fast they diffuse (heavy water boils at 101.4 °C and is about 11% denser than ordinary water).
- Some isotopes are stable; others are **radioactive**, their nuclei breaking down over time (carbon-14, tritium, uranium-235).
- They have the same place in the periodic table, which is where the name comes from: Greek *isos*, “same”, and *topos*, “place”.

Uses of isotopes:

- **Carbon-14** dates old wood, bone and cloth up to about 50,000 years old (radiocarbon dating), from how much of it is left.
- **Cobalt-60** gives out gamma rays used to kill cancer cells and to sterilize medical equipment.
- **Iodine-131** is used to find and treat diseases of the thyroid gland, which collects iodine.
- **Uranium-235** is the fuel of nuclear power stations: only 0.7% of natural uranium.
- **Deuterium**, as heavy water, slows down neutrons in some nuclear reactors.

Average atomic mass

Most elements are a mixture of isotopes, always in very nearly the same proportions wherever the element is found. Chlorine is about 75.76% chlorine-35 and 24.24% chlorine-37, whether it comes from sea salt or a rock. The element's atomic mass is the **average mass of its atoms, weighted by how common each isotope is**:

$$\text{average atomic mass} = \frac{(\text{mass}_1 \times \%_1) + (\text{mass}_2 \times \%_2) + \dots}{100}$$

WORKED EXAMPLE

Why chlorine is 35.45

Chlorine is 75.76% chlorine-35 (mass 34.969 u) and 24.24% chlorine-37 (mass 36.966 u). Find its average atomic mass. What answer do you get using the mass numbers, 35 and 37?

Answer

$$\text{average} = \frac{34.969 \times 75.76 + 36.966 \times 24.24}{100} = \frac{2649.25 + 896.06}{100} = 35.45$$

35.45, the atomic mass in every periodic table. With mass numbers:

$$\frac{35 \times 75.76 + 37 \times 24.24}{100} = \frac{2651.6 + 896.9}{100} = 35.48$$

about 35.5, the value often used in Nepal's exams. No single chlorine atom has a mass of 35.45: it is the average of a crowd in which about three atoms in four are chlorine-35.

WORKED EXAMPLE

Working backward

The atomic mass of gallium is 69.72. It has two isotopes, gallium-69 (mass 68.93 u) and gallium-71 (mass 70.92 u). What percentage of gallium atoms are gallium-69?

Answer

Let x % be gallium-69; then $(100 - x)$ % is gallium-71.

$$\frac{68.93x + 70.92(100 - x)}{100} = 69.72$$

$$68.93x + 7092 - 70.92x = 6972 \quad \Rightarrow \quad 1.99x = 120 \quad \Rightarrow \quad x = 60.3$$

About 60% gallium-69 and 40% gallium-71. (The measured share is 60.1%; the small difference comes from rounding the masses.)

Check 7

Two atoms have 17 protons each; one has 18 neutrons and the other 20. They are:

a isobars
b isotopes
c isotones
d atoms of two different elements

Isobars

Isobars are atoms of different elements that have the same mass number but different atomic numbers. They have the same total number of nucleons, made up of different numbers of protons and neutrons. (Greek *baros*, "weight": the same weight.)

Isobar	Protons	Neutrons	Mass number
$^{40}_{18}\text{Ar}$	18	22	40
$^{40}_{19}\text{K}$	19	21	40
$^{40}_{20}\text{Ca}$	20	20	40

Carbon-14 and nitrogen-14 are isobars too. Being different elements, isobars have **different chemical properties**: argon is an unreactive gas, potassium a soft metal that bursts into flame in water, calcium a metal in bones and limestone. Their masses are nearly, but not exactly, equal.

Isotones

Isotones are atoms of different elements that have the same number of neutrons but different numbers of protons, so different atomic numbers and different mass numbers.

Isotone	Protons	Neutrons	Mass number
$^{14}_{6}\text{C}$	6	8	14
$^{15}_{7}\text{N}$	7	8	15
$^{16}_{8}\text{O}$	8	8	16

Other isotones: sodium-23 and magnesium-24 (12 neutrons each); potassium-39 and calcium-40 (20 each).

A memory aid: isoto**p**es have the same number of **p**rotons, isoto**n**es the same number of **n**eutrons, and isob**a**rs the same m**a**ss number.

To find which group a pair belongs to, compare Z, N and A in turn:

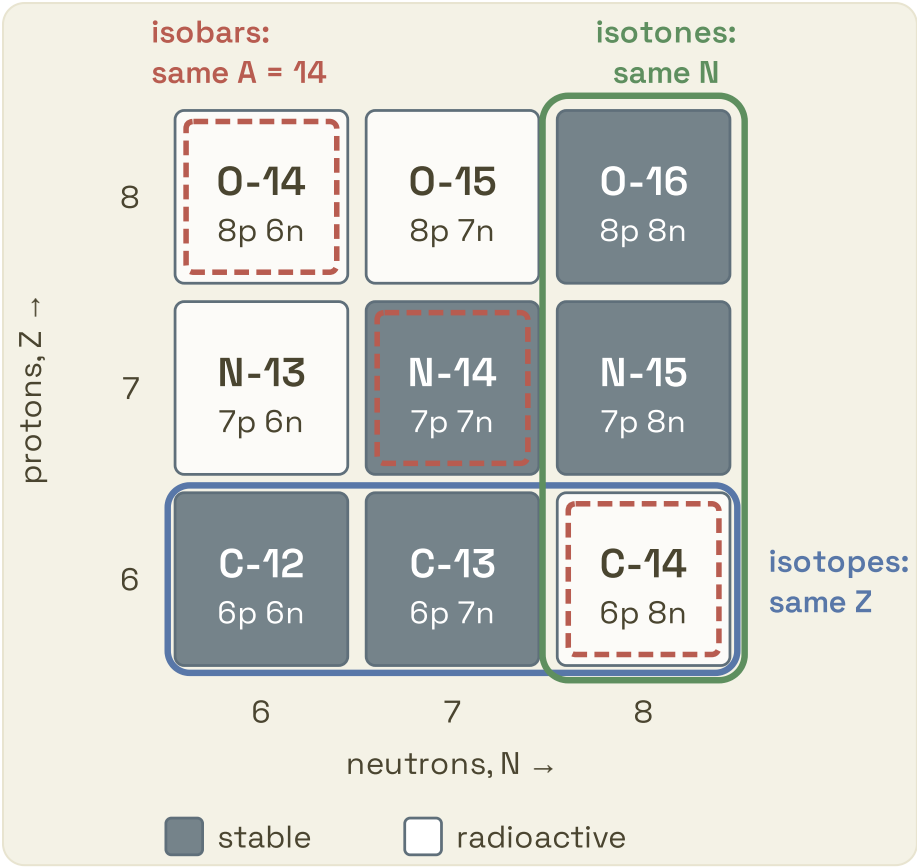


Figure 6.5 *Nine nuclei on a grid of protons (up) against neutrons (across). Isotopes lie along a row, isotones down a column, and isobars (ringed with dashes) along a diagonal: carbon-14 is in all three groups. The five shaded darker are stable; the other four are radioactive.*

	Atomic number, Z	Neutrons, N	Mass number, A	Same element?
Isotopes	same	different	different	yes
Isobars	different	different	same	no
Isotones	different	same	different	no

WORKED EXAMPLE

Sorting a list

From $^{14}_{6}\mathrm{C}$, $^{12}_{6}\mathrm{C}$, $^{14}_{7}\mathrm{N}$, $^{16}_{8}\mathrm{O}$, pick out a pair of isotopes, a pair of isobars and a pair of isotones.

Answer

First write $N = A - Z$ for each: C-14 has 8, C-12 has 6, N-14 has 7, O-16 has 8.

- **Isotopes** (same Z): C-14 and C-12.
- **Isobars** (same A): C-14 and N-14.
- **Isotones** (same N): C-14 and O-16, both with 8 neutrons.

Check 8

Which pair are isotones?

a $^{40}_{18}\mathrm{Ar}$ and $^{40}_{20}\mathrm{Ca}$

b $^{12}_{6}\mathrm{C}$ and $^{13}_{6}\mathrm{C}$

c $^{23}_{11}\mathrm{Na}$ and $^{24}_{12}\mathrm{Mg}$

d $^{1}_{1}\mathrm{H}$ and $^{2}_{1}\mathrm{H}$

IN SHORT

- **Dalton's atomic theory** (1803–1808): matter is made of indivisible atoms, the smallest unit of an element; atoms of one element are alike, of different elements different; they combine in simple whole-number ratios; in reactions they are only rearranged, never created or destroyed.
- **Modified:** atoms are made of electrons, protons and neutrons; isotopes have different masses; isobars of different elements have the same mass number; nuclear reactions change one element into another; many compounds have complex ratios. The atom is still the smallest unit of chemical reactions.
- **Electron** (Thomson, 1897, from cathode rays): charge -1.602×10^{-19} C, mass 9.109×10^{-31} kg (1/1836 of a proton's mass). **Proton** (canal rays, Goldstein 1886; named 1920): $+1.602 \times 10^{-19}$ C, 1.007 u. **Neutron** (Chadwick, 1932): no charge, 1.009 u.
- **Rutherford** (gold foil, 1909–1911): most of the atom is empty space; its positive charge and almost all its mass are in a tiny nucleus; electrons orbit it. It can't explain why atoms are stable or their line spectra.
- **Bohr** (1913): electrons move in fixed orbits of fixed energy without radiating; $mvr = \frac{nh}{2\pi}$; a jump between orbits gives out or takes in $\Delta E = E_2 - E_1 = h\nu$. For hydrogen, $E_n = -\frac{13.6}{n^2}$ eV.
- **Atomic number** Z = protons; **mass number** A = protons + neutrons; neutrons N = A − Z.
- **Atomic mass**: the average mass of an element's atoms compared with 1/12 of a carbon-12 atom; average atomic mass = Σ (isotope's mass × its %) ÷ 100.
- **Isotopes**: same Z, different A. **Isobars**: same A, different Z. **Isotones**: same N, different Z.

Exercise

1. Choose the correct option

a. Which part of Dalton's atomic theory is still accepted without change?

a atoms are indivisible

b all atoms of an element have the same mass

c an atom is the smallest particle that takes part in a chemical reaction

d atoms of different elements always have different masses

b. The particle discovered by James Chadwick in 1932 is the:

- **a** electron
- **b** proton
- **c** neutron
- **d** alpha particle

c. In the gold foil experiment, the fact that a very few alpha particles bounced back showed that:

- **a** electrons are very light
- **b** the positive charge and mass of an atom are concentrated in a tiny nucleus
- **c** most of the atom is empty space
- **d** gold atoms are negatively charged

d. According to Bohr, the angular momentum of an electron in an allowed orbit is:

- **a** nh
- **b** $\frac{nh}{2\pi}$
- **c** $\frac{h}{2\pi n}$
- **d** $\frac{2\pi}{nh}$

e. $^{40}_{18}\text{Ar}$ and $^{40}_{20}\text{Ca}$ are:

- **a** isotopes
- **b** isobars
- **c** isotones
- **d** the same element

2. Give reasons

a. An atom is electrically neutral.

b. Almost all the mass of an atom is in its nucleus.

c. Isotopes of an element have the same chemical properties.

3. Differentiate between

i. Atomic number and mass number

ii. Rutherford's atomic model and Bohr's atomic model

4. Answer the following questions

a. Write the main postulates of Dalton's atomic theory. How has each of the following been modified: (i) atoms are indivisible, (ii) atoms of an element are identical in mass?

b. How was the electron discovered? Write any four properties of cathode rays.

c. Describe Rutherford's alpha-ray scattering experiment, with a labeled diagram. Write its observations and conclusions.

d. Write the postulates of Bohr's atomic model. Give one merit and one limitation.

e. What are isotopes? Write their properties and any three uses.

5. Numerical problems

a. Find the numbers of protons, neutrons and electrons in (i) $^{27}_{13}\mathrm{Al}$, (ii) $^{56}_{26}\mathrm{Fe}$, (iii) $^{24}_{12}\mathrm{Mg}^{2+}$ and (iv) $^{16}_{8}\mathrm{O}^{2-}$.

b. An atom has 15 protons and 16 neutrons. Find its atomic number and mass number, write its symbol with them, and name the element.

c. Boron is 19.9% boron-10 (mass 10.013 u) and 80.1% boron-11 (mass 11.009 u). Find the average atomic mass of boron.

d. Copper has an atomic mass of 63.55 and two isotopes, copper-63 (mass 62.93 u) and copper-65 (mass 64.93 u). Find the percentage of each.

e. **Think:** A student says, “Since potassium-40 and calcium-40 have the same mass number, they must behave the same way in reactions.” Explain the mistake.

Found a mistake? Please tell us.