FROM QUANTA TO QUARKS

Mind map

CHAPTER 14

The models of the atom

Problems with the Rutherford model of the atom led to the search for a model that would better explain the observed phenomena

14.1 The early models of the atom

The most familiar model of the atom is one that involves many negatively charged electrons revolving around a central region known as the nucleus. The nucleus contains positively charged protons and neutral neutrons. This chapter will show that it took scientists many centuries to propose, debate, investigate and modify that familiar model of the atom (known as Bohr's model). The models of Democritus, Thomson, Rutherford and Bohr will be discussed in chronological order. These models are not the end of the development. New and more complicated models of the atom are constantly being developed. This option module will discuss this process.

Democritus

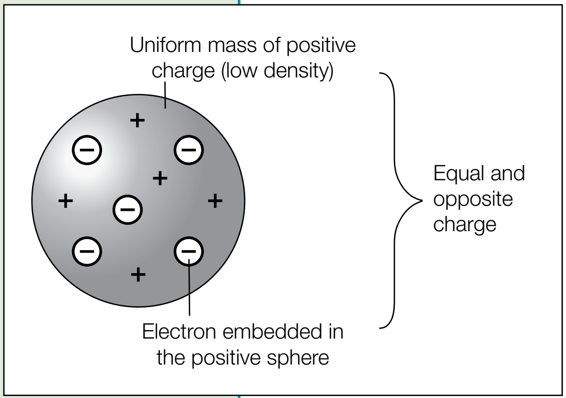
Democritus (c. 460–370 BC) was one of the earliest scientists to propose the model of the atom. He realised that if one kept dividing a substance into smaller and smaller pieces, there would be a point where further divisions could no longer be possible because one had reached the fundamental units that formed the substance. He proposed that these fundamental units were spherical in shape and were termed atoms. Therefore, Democritus proposed that atoms were indivisible particles that made up all matter.

Thomson's 'plum pudding' model

Thomson's experiment of the charge-to-mass ratio of cathode rays (1897) effectively indicated that, with any type of metal used as the cathode, identical cathode rays were obtained. Knowing that cathode rays were negatively charged particles—electrons—he proposed that atoms must all contain these common particles. The puzzle was, if all atoms were to contain electrons, how should these electrons be arranged inside the atoms?

Thomson developed his model where the atom was still assumed to be spherical in shape. However, this time the electrons were proposed to embed and scatter randomly among the region of the atom. Since the atom needed to be neutral overall, there must be positive charge to balance the negative electron charges. He proposed that the rest of the atom was uniformly positively charged, with its mass evenly distributed but low in density. This model was analogous to a plum pudding, where the electrons were like plums scattered throughout the 'pudding-like' atom, as shown in Figure 14.1.

Figure 14.1 Thomson's 'plum pudding' model of the atom



Rutherford's model of the atom

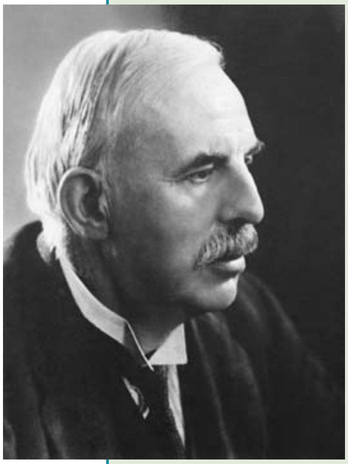
14.2

■ *Discuss the structure of the Rutherford model of the atom, the existence of the nucleus and electron orbits*

Rutherford's alpha particle scattering experiment

In 1911, Ernest Rutherford (1871–1937), or more precisely his students Geiger and Marsden, set out to perform an experiment that aimed to confirm Thomson's model of the atom, using the newly discovered alpha particles. Thomson's model showed electrons only occupying a very small space and the rest of the atom occupied with very low density positive charge. Rutherford thought that if this was the case, if alpha particles were fired at these atoms, they should either go straight through or through with very minimal deflections because nothing was in their way). He set up the experiment as shown in Figure 14.2.

However, the results of the experiment were surprising. Although most of the alpha particles went through the atoms with either no deflection or very small deflections as predicted, *one in eight thousand* alpha particles were deflected back at an angle greater than 90° (see Fig. 14.2). This was totally surprising, as it suggested that there must exist a sufficiently dense positively charged mass inside the atoms to cause the alpha particles to rebound. Repetition of the experiment achieved the same result.



Ernest Rutherford

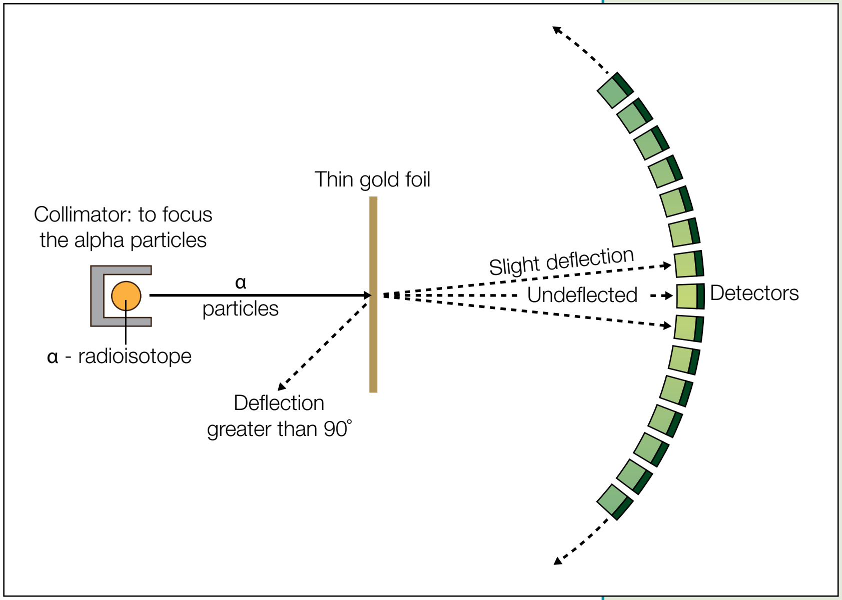


Figure 14.2 Rutherford's alpha particle scattering experiment

Rutherford's model of the atom

From the analysis of the results of his experiments, Rutherford proposed that the model of the atom needed to be modified to account for his observations. He stated that the only way the alpha particles could be deflected through such a large angle was if all of the atom's positive charge and nearly all of its mass was concentrated in a very small region, which he later named the **nucleus**. The electrons, (first proposed by Thomson), were to be placed around the nucleus in a circular fashion, and the rest of the atom consisted of empty space (see Fig. 14.3).

This model was adequate in explaining the deflection of the alpha particles. Usually these alpha particles would actually pass through the empty space between the nucleus and electrons, and hence would not have their path altered. If the alpha particles skimmed past the nucleus or collided with the electrons,

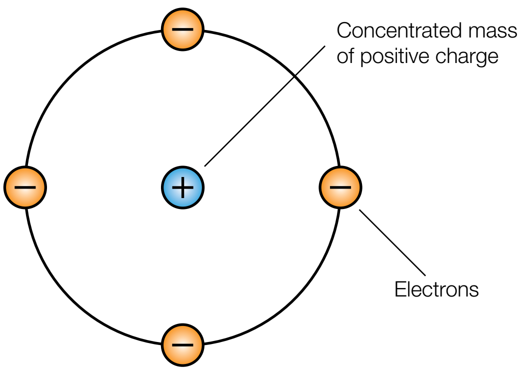


Figure 14.3 Rutherford's model of the atom

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Interactive Rutherford model

then their path would be altered slightly. Those that were deflected back at an angle greater than 90° must be due to the alpha particles colliding head on with the positively charged nucleus. However, since the nucleus was proposed to be very small compared to the size of the atom, the chance of this happening was remote.

PFA

H1

PFA 1: History of Physics:
'Evaluates how major advances in scientific understanding and technology have changed the direction or nature of scientific thinking'

■ *Discuss the structure of the Rutherford model of the atom, the existence of the nucleus and electron orbits*

How did Rutherford's proposed model of the atom differ from any before it?

The concept of the indivisible, structure-less model of the atom, as proposed by the chemist John Dalton (the 'billiard ball' model) was the accepted view of the atom for over 50 years in the 1800s. Dalton could not envisage 'empty space'; his atomic theory stated that atoms occupy all of the space in matter.

J. J. Thomson's work in 1897 suggested that the atom may indeed be divisible, as electrons were thought to be a part of any atom, and Goldstein subsequently showed in 1886 that atoms have positive charges. Rutherford's challenge was to devise a way of probing the atom in an attempt to find its structure. Thomson had proposed a 'plum pudding' model of the atom in which the negative charges, electrons, were embedded in a sea of positive fluid in the same way plums were embedded in a plum pudding. Lenard proposed yet another model in which positive and negative pairs were found throughout the atom.

Rutherford's experimental results, in which some alpha particles actually rebounded back off the thin gold foil and his careful analysis of this, led him to propose his 'planetary' model.

Why was this contribution a major advance in scientific understanding?

Rutherford's model of the atom was the first to propose a nucleus with the electrons in separate motion. The position of the electrons enabled advances in the field of chemistry, which deals with the interaction between the electrons of different atoms. There were problems with his model, however. Orbiting electrons should radiate electromagnetic radiation, lose energy and spiral into the nucleus, destroying the atom. Clearly, this did not happen.

How did it change the direction or nature of scientific thinking?

The motion of the electrons in Rutherford's model of the atom violated the laws of classical physics. However, rather than being disregarded, Rutherford's model triggered the further work of Bohr and others on their journey to develop quantum physics. The first step along this journey was to suggest that Rutherford's electrons could exist in a stable state and not emit radiation.

Evaluation of Rutherford's advance in scientific understanding and how it changed the nature and direction of scientific thinking

Rutherford's work paved the way for major changes in scientific thinking—the proposals of electrons orbiting a positive nucleus, and that much of the volume of atoms was empty space. The answers to Rutherford's puzzles led to the development of quantum theory and changes to the way in which matter was explained.

WWW

USEFUL WEBSITES

Background information on atomic theory around the time of Rutherford's experiment:
http://www.visionlearning.com/library/module_viewer.php?mid=50

A wealth of information on Ernest Rutherford (he was born in New Zealand):
http://www.rutherford.org.nz/

Inadequacies of Rutherford's model of the atom

Rutherford's model was quite successful in accounting for the surprising results of his experiment. However, there were still a few aspects that he was unable to explain.

- First, he could not explain the composition of what he called the nucleus. Although he said that most of the atom's mass and positive charges were to be concentrated into this very small and dense area called the nucleus, he could not explain what was in the nucleus. (The existence of protons and neutrons were not known at the time).
- Although he proposed that the electrons should be placed around the nucleus, he did not know how exactly to arrange the electrons around the nucleus, except 'like planets around the Sun'.
- The biggest problem that Rutherford failed to explain was how the negative electrons could stay away from the positive nucleus without collapsing into it. The only way to overcome the attractive force between the positive nucleus and the negative electrons was to have the electrons orbiting around the nucleus, much like the Moon going around the Earth. However, electrons, when circulating around the nucleus, would have centripetal acceleration (centripetal acceleration applies to all circular motion). *Accelerating charges produce EMR.* This meant that electrons would release EMR as they were orbiting and these EMR would radiate away, which posed a loss of energy. This loss of energy must be derived from the kinetic energy of the electrons (law of conservation of energy), resulting in the electrons slowing down. Eventually, the electrons would lose enough kinetic energy so that they would no longer have sufficient velocity to maintain the orbit around the nucleus and would spiral back into the nucleus. Obviously, this did not happen—but Rutherford's model failed to provide a reason for this.

14.3 Planck's hypothesis

- ***Discuss Planck's contribution to the concept of quantised energy***

The concept of quantisation of energy has been discussed in Chapter 11. The basic definition is as follows.

Definition

The radiation emitted from a black body is not continuous as waves; it is emitted as packets of energy called **quanta** (photons).

The energy of each of these quanta or photons is related to their frequency by the equation:

$$E = hf$$

Where:

E = the energy of each quantum or photon, measured in J

h = Planck's constant, which has a value of 6.626×10^{-34} J s

f = the frequency of the radiation (wave), measured in Hz

Planck's hypothesis was initially made in order to theoretically derive the black body radiation curve. However, it was later used by Einstein to successfully explain the photoelectric effect. In this module, you will see this hypothesis also forms the basic foundation for the quantum theory and quantum mechanics. Also, as you will see later in this chapter, Planck's hypothesis forms an essential part of Niels Bohr's model of the atom.

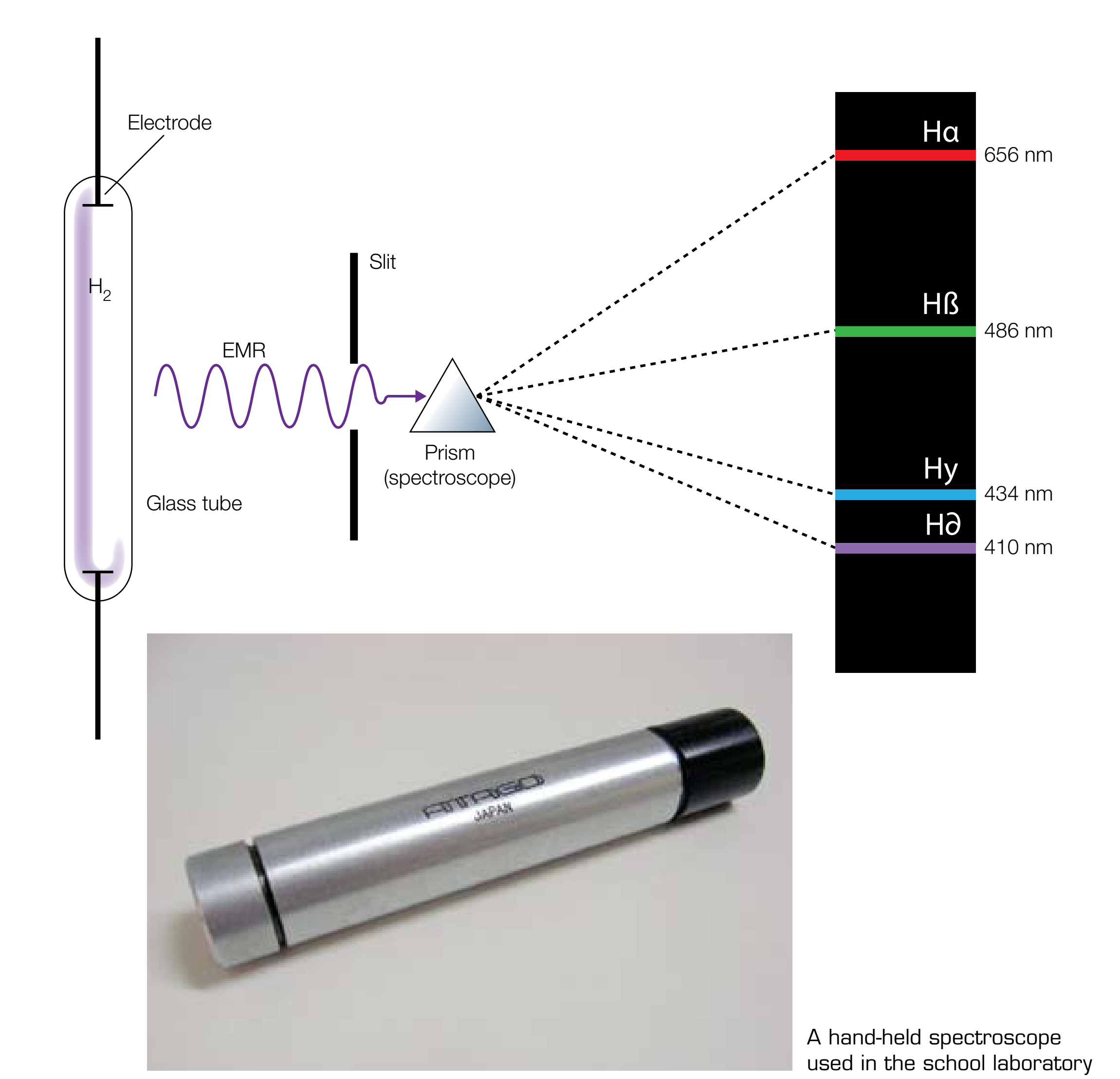
14.4 The hydrogen emission spectrum

A neon light

Using a neon light as an everyday example: A neon light consists of a glass tube containing neon gas in which two electrodes are embedded. When electricity is passed through the tube, the gas glows to produce the 'neon light'. This does not just happen with the neon element, rather it happens with other elements in the periodic table. Our discussion will focus on the hydrogen element.

As shown in Figure 14.4, a glass tube with two electrodes contains hydrogen gas. When electricity is passed between the electrodes, the hydrogen gas glows purple-red. Now we need to introduce a new concept that this emitted light is not just a single colour (hence a single wavelength), but rather a combination of different wavelengths of light. When this light is separated into its individual wavelengths, for example by a prism in a device called a **spectroscope**, and it is cast onto a black background, one can clearly see these individual colours, and hence wavelengths (see Fig. 14.4). This light pattern is known as the **hydrogen emission spectrum**. As you can see in Figure 14.4, the hydrogen emission spectrum has a pattern that consists of red and blue lines at different wavelengths. A similar method can be used to obtain the emission spectrum for other elements. It is important to realise that the emission spectrum is unique to an element, such that each element has its own unique pattern of wavelengths. Consequently, these wavelengths may also be used to identify any unknown element.

Another model of the atom is needed to understand the mechanism of the production of the emission spectra—Bohr's model.



A hand-held spectroscope used in the school laboratory

Figure 14.4 Hydrogen emission spectrum

14.5 Bohr's model of the atom: introduction

- ***Define Bohr's postulates***
- ***Analyse the significance of the hydrogen spectrum in the development of Bohr's model of the atom***

In 1913, Danish physicist Niels Bohr (1885–1962) developed a model of the atom. This model was based on the ideas of the quantisation of energy and Planck's hypothesis, Rutherford's model of the atom, as well as by observing the pattern of the hydrogen emission spectrum. In simple terms, Bohr's model was based on three fundamental postulates.

Postulate 1: All electrons around the nucleus are only allowed to occupy certain fixed positions and energy levels outward from the nucleus, thus the electron orbits are quantised and are known as the principal energy shells. While in a particular orbit, electrons are in a stationary state and do not radiate energy.

Postulate 2: When an electron moves from a lower orbit to a higher orbit, or falls down from a higher orbit to a lower orbit, it will absorb or release a quantum of energy (EMR). The energy of the quantum is related to the frequency of the EMR by the formula: $E = hf$, where E is the energy of the quantum (J), h is the Planck's constant and f is the frequency of the EMR (Hz).

Postulate 3: The electrons' angular momentum is quantised as mvr_n (angular momentum) $= \frac{nh}{2\pi}$, where n is the principal energy shell number, and the most inner energy shell is assigned n = 1.

NOTE: Postulate 3 was a purely empirical formula derived from the measurements taken from the hydrogen emission spectrum.

Based on his first postulate, Bohr modified Rutherford's model to one that contained a central positively charged nucleus and many orbits for the electrons around the nucleus, as shown in Figure 14.5. As mentioned, he called these orbits the **energy shells** or the **principal energy shells**. Electrons, when they were occupying these shells, were said to be stable, and did not need to rotate in order to stay away from the nucleus, nor did they radiate energy.

There are also mathematical consequences to these postulates. Based on his first postulate and third postulate, combining with equations from classical physics—centripetal force and Coulomb's law, Bohr was able to develop a series of mathematical equations to describe quantitatively the radius as well as the energy change of the principal energy shells as they moved away from the nucleus.

$$r_n = \frac{h^2 n^2}{4\pi^2 kme^2}$$

$$E_n = \frac{-2\pi^2 k^2 me^4}{h^2 n^2}$$

Where:

r_n = radius of the nth principal energy shell (m)

E_n = energy of the nth principal energy shell (J)

h = Planck's constant, 6.626×10^{-34} J s

n = principal energy shell number

k = constant, 9.11×10^{-31}

e = charge of electron, 1.602×10^{-19} C

m = mass of electron, 9.109×10^{-31} kg

Also the equations can be re-written with reference to the radius and energy of the first principal energy shell, that is, r_1 and E_1.

$$r_n = n^2 r_1$$

$$E_n = \frac{1}{n^2} E_1$$

Where:

$$r_1 = \frac{h^2 1^2}{4\pi^2 kme^2} = \frac{h^2}{4\pi^2 kme^2}$$

$$E_1 = \frac{-2\pi^2 k^2 me^4}{h^2 1^2} = \frac{-2\pi^2 k^2 me^4}{h^2}$$

Note that both E_1 and r_1 are constants.

Furthermore, the transition between the electron orbits, such as mentioned in postulate 2, explains the mechanism by which the hydrogen emission spectrum is produced. This will be discussed in detail in the next section.

Bohr's model and the hydrogen emission spectrum

14.6

- ***Describe how Bohr's postulates led to the development of a mathematical model to account for the existence of the hydrogen spectrum:***

$$\frac{1}{\lambda} = R\left(\frac{1}{n_f^{\,2}} - \frac{1}{n_i^{\,2}}\right)$$

According to Bohr, when electrons absorb energy, they will move up to a higher orbit. Their energy may be given by the means of heat, or electricity through the electrodes embedded in the glass tube as shown in Figure 14.4. When this energy is withdrawn, the excited electrons will not to stay in these higher orbits all the time. They later fall back to lower orbits. As the second postulate states, when the electrons fall back to lower orbits, they radiate energy in the form of **EMR**, the frequency of which is directly proportional to the difference in energy between the two levels ($E = hf$). This is illustrated in Figure 14.5.

$$\Delta E = E_i - E_f = hf$$

Where:

ΔE = change in energy as the electron transits from one orbit to another (J)

h = Planck's constant, 6.626×10^{-34} J s

f = frequency of the EMR released (Hz)

E_i = the energy of the initial orbit, that is, where electron falls from (J)

E_f = the energy of the final orbit, that is, where electron falls to (J)

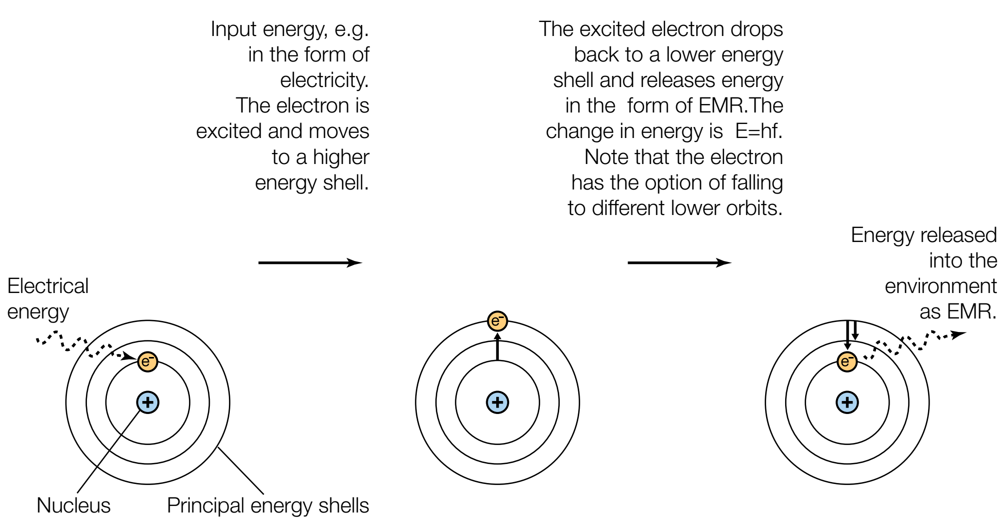


Figure 14.5 The formation of the hydrogen emission spectrum

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Simulation: Bohr's theory of the hydrogen atom

It is important to emphasise that when electrons fall back, they can go straight from the excited state to the lowest possible energy level or do this by many discrete steps. For example, an electron that is to fall from the fifth shell (n = 5) to the 1st shell (n = 1) has the choice of going from 5 to 1 directly, or from n5 to n4 and then one shell at a time, or from n5 to n3 then to n1, or many other combinations. Basically, the falling of electrons is a probability function. These combinations lead to different energy changes, and hence the different wavelengths and therefore colours seen in the emission spectrum (see Fig. 14.5).

To further add to the equation above, we also know that the energy level of each shell can be described by the equation:

$$E_n = \frac{1}{n^2}E_1$$

Hence we have:

$$\Delta \mathrm{E} = \mathrm{hf} = E_i - E_f$$

$$\mathrm{hf} = \frac{1}{(n_i)^2}E_1 - \frac{1}{(n_f)^2}E_1$$

as $c = f\lambda$

$f = \dfrac{c}{\lambda}$ and which can be substituted into the above equation:

$$h\frac{c}{\lambda} = E_1\left[\frac{1}{(n_i)^2} - \frac{1}{(n_f)^2}\right]$$

$$h\frac{c}{\lambda} = -E_1\left[\frac{1}{(n_f)^2} - \frac{1}{(n_i)^2}\right]$$

$$\frac{1}{\lambda} = \frac{-E_1}{hc}\left[\frac{1}{(n_f)^2} - \frac{1}{(n_i)^2}\right]$$

$\dfrac{1}{\lambda} = R\left[\dfrac{1}{(n_f)^2} - \dfrac{1}{(n_i)^2}\right]$, where R is $\dfrac{-E_1}{hc}$ and is called **Rydberg's constant**.

Since E_1 is negative, R is positive. Therefore:

$$\frac{1}{\lambda} = R\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

Where:

λ = wavelength (m)

R = Rydberg's constant, $1.097 \times 10^7\ \mathrm{m^{-1}}$

n_f = principal energy shell the electron falls to

n_i = principal energy shell the electron falls from

This equation is helpful as it can be used to calculate the frequencies and wavelengths of the emission spectrum provided that we know the initial and final orbit of the transition. Therefore, it is clear that Bohr's model essentially provides a **theoretical explanation** for the appearance of the hydrogen emission spectrum.

■ *Solve problems and analyse information using:*

$$\frac{1}{\lambda} = R\left(\frac{1}{n_f^{\,2}} - \frac{1}{n_i^{\,2}}\right)$$

Some examples:

Example 1

What is the wavelength of the emission line when an electron falls from the third energy shell to the second energy shell? Describe the nature of this EMR.

Solution

Using $\frac{1}{\lambda} = R\left[\frac{1}{(n_f)^2} - \frac{1}{(n_i)^2}\right]$,

Where $R = 1.097 \times 10^7$

$n_f = 2$

$n_i = 3$

$\lambda = ?$

$$\frac{1}{\lambda} = 1.097 \times 10^7 \times \left[\frac{1}{2^2} - \frac{1}{3^2}\right]$$

$$= 6.56 \times 10^{-7}\ \text{m}$$

The wavelength is calculated to be 6.56×10^{-7} m. This emission spectrum line corresponds to the first red light of the hydrogen spectrum. Note that similar calculations may apply to the other wavelengths.

Example 2

Calculate the lowest possible frequency of an emission line, if the final energy shell is always 1. Describe the nature of the radiation.

Solution

The lowest possible frequency corresponds to the smallest energy difference during orbital transition. Hence if the final orbit is 1, then the initial orbit can only be 2 (to give the smallest difference). Hence:

Using $\frac{1}{\lambda} = R\left[\frac{1}{(n_f)^2} - \frac{1}{(n_i)^2}\right]$,

Where $R = 1.097 \times 10^7$

$n_f = 1$

$n_i = 2$

$\lambda = ?$

$$\frac{1}{\lambda} = 1.097 \times 10^7 \times \left[\frac{1}{1^2} - \frac{1}{2^2}\right]$$

$$\lambda \approx 1.22 \times 10^{-7}\ \text{m}$$

Then $c = f\lambda$

$$f = \frac{c}{\lambda}$$

$$f = \frac{3 \times 10^8}{1.22 \times 10^{-7}}$$

$$f \approx 2.47 \times 10^{15}\ \text{Hz}$$

This frequency is in the UV range.

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Worked example 25

Example 3

An emission spectrum line has the wavelength of 1.88×10^{-6} m (infrared). Suppose the final orbit is 3, calculate the orbit from which the electron falls from.

Solution

Using $\frac{1}{\lambda} = R\left[\frac{1}{(n_f)^2} - \frac{1}{(n_i)^2}\right]$,

Where $R = 1.097 \times 10^7$

$n_f = 3$

$\lambda = 1.88 \times 10^{-6}$

$n_i = ?$

$$\frac{1}{1.88 \times 10^{-6}} = 1.097 \times 10^7 \times \left[\frac{1}{3^2} - \frac{1}{n_i^2}\right]$$

$$\left[\frac{1}{9} - \frac{1}{n_i^2}\right] = \frac{1}{1.88 \times 10^{-6} \times 1.097 \times 10^7}$$

$$\frac{1}{n_i^2} = \frac{1}{9} - \frac{1}{1.88 \times 10^{-6} \times 1.097 \times 10^7}$$

$$\frac{1}{n_i^2} = 0.0626$$

$$n_i^2 = 15.97$$

$$n_i = 4$$

Balmer series

From the previous examples, an emission spectrum can also include ultra-violet and infrared in addition to visible light. The visible parts of the spectral lines were observed first. Years before Bohr had proposed his theory to account for the hydrogen emission spectrum, a Swiss school teacher, Johan Balmer (1825–1898), in 1885 realised that the visible part of the hydrogen emission spectrum obeyed a simple mathematical relationship, such that:

$$\lambda = b\left(\frac{n^2}{n^2 - 2^2}\right)$$

However, Balmer was only able to derive this equation using pure mathematics based on observational data and did not have any reasons for why this equation was so. Such equations are known as empirical. The visible part of the hydrogen spectrum was named **Balmer series** in his honour.

Bohr realised that the '2^2' in the equation was due to the fact that all emission lines from the visible part of the hydrogen emission spectrum were the result of electrons falling to the second energy shell ($n_f = 2$). Hence Bohr's 'modern formula' for the Balmer series is:

$$\frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{n_i^2}\right)$$

Bohr's model and his postulates enabled scientists to account for the hydrogen emission spectrum as well as allowing them to calculate the wavelengths on a theoretical basis. He provided a theoretical explanation for Balmer's formula, which was otherwise derived from the empirical observations. Therefore, his model of the atom was quite successful overall, not only structurally but also functionally.

Other spectral line series have been observed. For instance, when electrons fall back to the first orbit, the EMR released is all in the ultra-violet range, and the series is called **Lyman series**. When electrons are excited and fall back to the third orbit, the EMR released is always in the infrared range, and is called **Paschen series**. There are many more, corresponding to the final energy level the electrons fall to.

More on the hydrogen emission spectrum

14.7

■ *Process and present diagrammatic information to illustrate Bohr's explanation of the Balmer series*

To summarise the energy profile of the electron shells around an atom and the different types of spectral series mentioned above, we use a diagram (see Fig. 14.6). Note that the energy differences between the orbits form a converging pattern; that is, as the principal energy shell goes outwards from the nucleus, the difference in energy between each energy shell gets smaller. For example, the energy difference between n = 1 and n = 2 is large compared to that between n = 5 to n = infinity (see Fig. 14.6).

Would you be able to work out the reason for this based on the equation,

$E_n = \frac{1}{n^2}E_1$?

The consequence of this is that the type of radiation emitted is determined by where the electron finally falls to. This is because the lower orbits have larger energy gaps. With more contribution to the energy difference, they have more weight in determining the type of radiation. For instance, when an electron falls to the first shell, the energy difference between n = 2 and n = 1 is so large that regardless of where the electron falls from, the energy difference is going to be large enough to cause the emission to be in the ultra-violet range. When an electron falls to the third

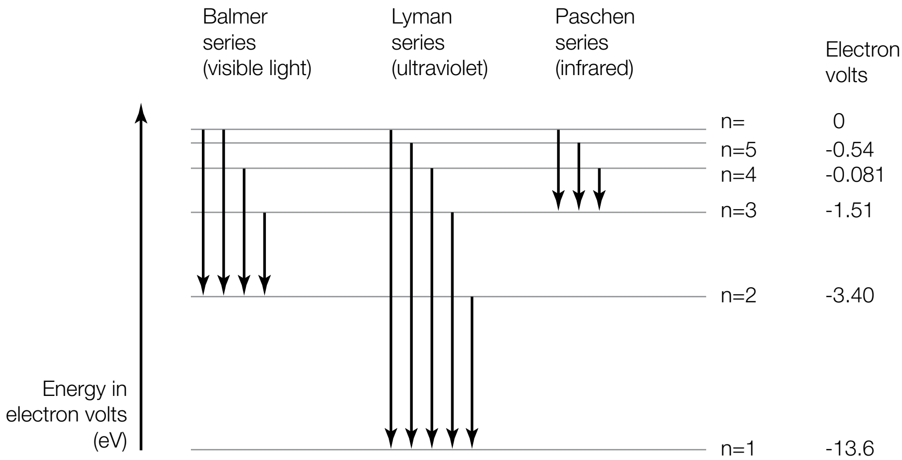

Figure 14.6 A summary of the hydrogen emission spectrum

shell, the energy gap is always going to be small, therefore the emission is always going to be in the infrared range.

Furthermore, the energy differences between emission lines on the red side of the spectrum are always larger than those on the blue side of the spectrum. Consequently, the differences in frequency thus wavelength for emission lines on the red side of the spectrum will be larger than those on the blue side of the emission spectrum. With any atomic emission spectrum, the lines towards the red side are further apart than the lines towards the blue side. Refer to Figure 14.4 again.

14.8 Limitations of Bohr's model of the atom

- ***Discuss the limitations of the Bohr model of the hydrogen atom***
- ***Analyse secondary information to identify the difficulties with the Rutherford-Bohr model, including its inability to completely explain:***
  - ***the spectra of larger atoms***
  - ***the relative intensity of spectral lines***
  - ***the existence of hyperfine spectral lines***
  - ***the Zeeman effect***

Bohr's model was overall quite successful. First, it provided a reason (although without proof) why electrons were able to stay away from the nucleus. Second, it explained the hydrogen emission spectrum. However, there were still a few fundamental inadequacies, as discussed below:

1. Bohr's model used a mixture of classical physics and quantum physics without giving any reasons for that. The classical physics in his model included circular motion of the electrons, and the concept of angular momentum as well as Coulomb's law; all were used to derive the equations for the radius and energy of each energy shell. The quantum physics aspect of his model included the quantisation of the electron orbits and transitions, quantisation of energy, that is, $E = hf$ as well as the quantisation of angular momentum as $\frac{nh}{2\pi}$. Furthermore, his quantum physics theories were radical and lacked rational explanation. For instance, in his first postulate, Bohr stated an electron was to be stable when it was in its orbit; although this explained how an electron could stay away from the nucleus, it was not logically convincing enough. Also, he could not explain why the electron's angular moment was quantised as $\frac{nh}{2\pi}$.

   Historically, Bohr's model of the atom was known as the **quantum theory**, a hybrid between the classical Newtonian physics and a brand new area of physics called **quantum mechanics**, which we will discuss briefly in Chapter 15.
2. The model could not explain the **relative intensity** between spectral lines. It was observed in experiments that some spectral lines were more intense, that is, brighter than others, which indicated that some types of transitions were more preferred than others. Bohr's model failed to explain this.

3. The model did not work for multi-electron atoms. Bohr's mathematics and equations worked well and accurately for hydrogen atoms. However, they failed when Bohr tried to apply them to atoms with more than one electron, even helium atoms. Obviously, this was inadequate, as a good model should work for all types of atoms.
4. The model could not explain the existence of **hyperfine spectral lines**. Hyperfine spectral lines are thin, faint lines that exist as a cluster around a main spectral line (sometimes, they make up a main spectral line). They sit very close together and require close observation to distinguish between them. Bohr's model only allowed the prediction for the main spectral lines, but could not explain why some transitions were outside these main spectral lines, which caused the hyperfine lines.
5. Bohr's model could not explain the **Zeeman effect**. The Zeeman effect is defined as the splitting of the spectral lines when a powerful magnetic field is applied. Bohr's model could not give a satisfactory explanation for this phenomenon.

Furthermore, just like Rutherford's model, Bohr's model did not include an explanation for the structure of the nucleus. This was later done by other scientists, as discussed in Chapter 16.

You are encouraged to conduct your own research to further extend your understanding and appreciation of the relative intensity of spectral lines, the existence of hyperfine spectral lines and the Zeeman effect. Use these as key words to facilitate your research, using either Internet or library resources. Multimedia sources, such as videos or animations, will be particularly useful.

Observe the visible components of the hydrogen emission spectrum

■ ***Perform a first-hand investigation to observe the visible components of the hydrogen spectrum***

The basic principle of the apparatus used to obtain a hydrogen emission spectrum and the expected results were discussed earlier in this chapter. Familiarise yourself with the set-up and the underlying physics theories and appreciate the colours of the emitted light. You should also be familiar with the calculations and equations used to determine the wavelength of the spectral lines.

FIRST-HAND INVESTIGATION

PFAs

H1

PHYSICS SKILLS

H12.1A, B, D
H12.2B

CHAPTER REVISION QUESTIONS

1. Describe Thomson's 'plum-pudding' model of the atom.
2. In 1911, Rutherford's assistants, Geiger and Marsden, performed the now famous experiment using alpha particle scattering.
   (a) Describe the procedure of the experiment with the aid of a diagram.
   (b) What relevant observations were made?

(c) Did the result contradict the knowledge of atoms at the time?
(d) What was the conclusion drawn from the experiment? What was the implication of this conclusion?

3. Draw up a table to contrast the differences (three or more) between quantum physics and classical physics.

4. Explain how an atomic emission spectrum is produced and describe two features of an atomic emission spectrum.

5. (a) Define Bohr's three postulates in regard to the structure of the atom.
(b) How do these three postulates lay down the foundation for his model of the atom?

6. With the aid of equations, explain the pattern of change in radius and energy of the electron shells from the innermost one to the outermost.

7. What is the Balmer series and how is it related to Bohr's theory of the atom?

8. Define the Lyman and Paschen series.

9. (a) Calculate the wavelength of the EMR emission when an electron falls from the fifth shell to the third shell.
(b) Calculate the wavelength of the EMR emission if the orbital transition of the electron is from the sixth level to the second level.
(c) Calculate the energy required to raise an electron from the first energy level to the seventh energy level.
(d) Calculate the frequency of the EMR emitted when an electron falls from the third energy level to the second energy level; hence, determine the nature of the EMR.

10. Calculate the first ionisation energy of a sodium atom.

11. Evaluate the successes and inadequacies of Bohr's model of the atom.

SR

Answers to chapter revision questions