CHAPTER 14

The models of the atom

Answers to revision questions

1. Thomson's 'plum-pudding' model is such that the electrons are embedded in a low density positively charged sphere. The negative charges carried by the electrons are exactly balanced out by the positive sphere.

2. (a) See Figure 14.2.

   (b) Most of the alpha particles went through the thin gold foil either undeflected or with a slight deflection. However, 1 in 8000 alpha particles were deflected back at an angle greater than 90°.

   (c) Yes, since the model of the atoms at the time could not explain why alpha particles were bounced back.

   (d) The conclusion drawn was that the only way the alpha particles could be deflected through such a large angle was if all the atom's positive charges and nearly all of its mass were concentrated in a very small region, which was later termed the nucleus.

3.

Classical physics	**Quantum physics**
Quantities are seen as continuous	Quantities are seen as discrete
Traditional, has been around for hundreds of years	Relatively new, was only developed during the last few decades
More useful for macroscopic objects	More useful for microscopic objects

4. When electrons absorb energy, they will move up to a higher orbit. The energy may be given by the means of heat or electricity. When this energy is withdrawn, the excited electrons will not stay in these higher orbits but will fall back to lower orbits. 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. These transitions of electrons between energy levels result in the emission spectrum.

   The features of an atomic emission spectrum include:
   - unique to the element
   - have different colour lines
   - the lines on the red side are always further apart than the lines on the blue side
   - there exist hyperfine lines.

5. (a) *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 principle 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 $E = hf$.

   Postulate 3: The electrons' angular momentum is quantised as $mvr_n = \frac{nh}{2\pi}$.

(b) The first postulate gives rise to the energy shells as well as explaining the stability of the electrons in these energy shells.

The first and third postulates, along with other classical physics laws, give rise to the equations that describe the radius and energy of the electron shells.

The second postulate theoretically explains the hydrogen emission spectrum.

6. From $E_n = \frac{1}{n^2}E_1$, the energy of the electron shells forms a convergent pattern. Whereas from $r_n = n^2 r_1$, the radius of the energy shells forms a divergent pattern.

7. A Swiss school teacher, Johan Balmer, in 1885, realised that the visible part of the hydrogen emission spectrum obeyed a simple mathematical relationship, such that:

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

Hence the visible part of the emission spectrum was named the Balmer series in his honour.

Bohr theoretically explained the Balmer series and stated that the series could be obtained when electrons fall from higher energy shells to the second energy shell.

8. **Lyman series** refers to the UV part of the emission spectrum; it happens when the electron falls from higher energy shells back to the first shell.

Paschen series refers to the infrared part of the emission spectrum; it happens when the electron falls from higher energy shells back to the third shell.

9. (a) Using:

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

$$R = 1.097 \times 10^7 \text{ m}^{-1}$$

$$n_f = 3$$

$$n_i = 5$$

$$\lambda = ?$$

$$\frac{1}{\lambda} = 1.097 \times 10^7 \times \left(\frac{1}{3^2} - \frac{1}{5^2}\right)$$

$$\lambda = 1.28 \times 10^{-6} \text{ m}$$

(b) Using:

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

$$R = 1.097 \times 10^7 \text{ m}^{-1}$$

$$n_f = 2$$

$$n_i = 6$$

$$\lambda = ?$$

$$\frac{1}{\lambda} = 1.097 \times 10^7 \times \left(\frac{1}{2^2} - \frac{1}{6^2}\right)$$

$$\lambda = 4.10 \times 10^{-7} \text{ m}$$

(c) Using:

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

$R = 1.097 \times 10^7\ \text{m}^{-1}$

$n_f = 1$

$n_i = 7$

$\lambda = ?$

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

$\lambda = 9.31 \times 10^{-8}\ \text{m}$

$E = \frac{hc}{\lambda} (E = nf,\ \text{where } f = \frac{c}{\lambda})$

$h = 6.626 \times 10^{-34}$

$c = 3 \times 10^8\ \text{m s}^{-1}$

$\lambda = 9.31 \times 10^{-8}\ \text{m}$

$E = ?$

$$E = \frac{6.626 \times 10^{-34} \times 3 \times 10^8}{9.31 \times 10^{-8}}$$

$= 2.14 \times 10^{-18}\ \text{J}$

(d) Using:

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

$R = 1.097 \times 10^7\ \text{m}^{-1}$

$n_f = 2$

$n_i = 3$

$\lambda = ?$

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

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

$= 656\ \text{nm}$

$\therefore$ This is the wavelength of the first red light of the hydrogen spectrum.

10. The energy released when an electron falls from infinity to the third energy shell can be calculated as:

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

$R = 1.097 \times 10^7 \text{ m}^{-1}$

$n_f = 3$

$n_i = \infty$

$\lambda = ?$

$$\frac{1}{\lambda} = 1.097 \times 10^7 \times \left(\frac{1}{3^2} - \frac{1}{\infty^2}\right)$$

$$\left(\frac{1}{\infty^2} = 0\right)$$

$\lambda = 8.20 \times 10^{-7} \text{ m}$

$$E = \frac{hc}{\lambda}$$

$h = 6.626 \times 10^{-34}$

$c = 3 \times 10^8 \text{ m s}^{-1}$

$\lambda = 8.20 \times 10^{-7} \text{ m}$

$E = ?$

$$E = \frac{6.626 \times 10^{-34} \times 3 \times 10^8}{8.20 \times 10^{-7}}$$

$$= 2.42 \times 10^{-19} \text{ J}$$

∴ This is the same energy moving an electron from shell n = 3 to infinity, therefore the first ionisation energy of sodium is approximately 2.42×10^{-19} J.

11. Bohr's model was, overall, quite successful. First, it provided a reason why electrons were able to stay away from the nucleus; second, it explained the hydrogen emission spectrum. However, there were some fundamental inadequacies:

- Bohr's model used a mixture of classical physics and quantum physics without giving any reasons.
- The model could not explain the relative intensity between spectral lines.
- The model did not work for multi-electron atoms.
- The model could not explain the existence of hyperfine spectral lines.
- Bohr's model could not explain the Zeeman effect.

Furthermore, Bohr's model did not include a discussion on the structure of the nucleus.