Questions

Module 8: From the Universe to the Atom

8.3 Quantum Mechanical Nature of the Atom

Multiple-choice questions: 1 mark each

1. The signal from a microwave transmitter can be thought of as a beam of photons. The photons from a particular transmitter have a wavelength of 3.5×10^{-2} m.

   What is the approximate energy of each photon?

   (A) 7.73×10^{-44} J

   (B) 5.68×10^{-24} J

   (C) 2.32×10^{-35} J

   (D) 1.89×10^{-32} J

2001 HSC Q6

2. What is the energy of a photon of wavelength 580 nm?

   (A) 3.43×10^{-19} J

   (B) 3.43×10^{-28} J

   (C) 3.85×10^{-31} J

   (D) 3.85×10^{-40} J

2008 HSC Q13

3. What is the wavelength, in metres, of a photon with an energy of 3.5 eV?

   (A) 1.2×10^{-6}

   (B) 3.5×10^{-7}

   (C) 1.18×10^{-15}

   (D) 5.67×10^{-26}

2016 HSC Q11

4. Louis de Broglie proposed that electrons exhibited wave properties under suitable conditions and the wavelength (λ) of any particle with momentum could be calculated as: $\lambda = \frac{h}{mv}$.

What experimental evidence was able to verify this prediction?

(A) Davisson and Germer's diffraction experiment showing electrons had a wave nature.

(B) Davisson and Germer' diffraction experiment showing electrons had a particle nature.

(C) Schrödinger demonstrated the quantum mechanical nature of matter.

(D) The Rutherford-Bohr atomic model and the Balmer series in hydrogen.

5. A laser emits light of wavelength 550 nm. Calculate the frequency of this light.

(A) 18.3 Hz

(B) 5.45×10^{21} Hz

(C) 5.45×10^{5} Hz

(D) 5.45×10^{14} Hz

Adapted 2017 HSC Q21(a)

6. How did Schrödinger's model of the atom compare to the Bohr model of the atom?

(A) Schrödinger's model involved a thought experiment using a cat in a box, and Bohr's did not.

(B) Schrödinger had a quantum model with defined positions for the electron orbits, similar to Bohr's model.

(C) Schrödinger's model had electron probability clouds, instead of the electrons being in known positions.

(D) Both Schrödinger's and Bohr's models of the atom only applied to the hydrogen atom.

Short-answer questions

7. (a) Outline the key features of the Rutherford model of the atom.

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(b) Explain how the Bohr model of the atom overcomes an identified limitation of the Rutherford model of the atom.

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2013 HSC Q35(a) ... 2 + 3 = 5 marks

[Same Q's as in: 2003 HSC Q31(a)]

8. Identify TWO limitations of the Bohr model of the atom.

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2018 HSC Q34(a)(ii) ... 3 marks

This question is the same / similar to • 2004 HSC Q31(d)(iii)

• 2016 HSC Q34(b)(i) • 2012 HSC Q34(a)(ii) • 2007 Q31(a)(ii) • 2008 HSC Q31(c)

[Note: For example, another way this question was asked:

'Outline TWO features of the hydrogen spectrum that Bohr's model could not explain.']

9. State the major difficulty with Rutherford's model of the atom.

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1995 HSC Q32C(d)(iv) ... 2 marks

10. (a) Some students set up a hydrogen spectral tube in a darkened laboratory with an induction coil attached to it as a high voltage power source. When the emission spectrum of hydrogen was produced, they examined it.

What technology would they have used to observe the visible components of the hydrogen emission spectrum and what would they have seen?

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(b) How would the results from this investigation support Bohr's model of the atom?

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New part (a). Part (b) is adapted 2008 HSC Q31(a)(ii)... 3 + 2 = 5 marks

[Part (a) above is similar to part of: 2004 HSC Q31(d)(i)]

11. How does the Bohr model of the atom account for the emission spectrum of hydrogen?

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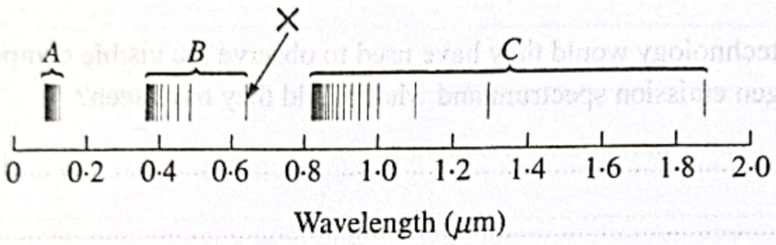
2018 HSC Q34(a)(ii)... 3 marks

12. How did Bohr overcome the stability issues posed by Rutherford's model of the atom?

Adapted from 1990 HSC Q Elective 1A (b)(ii) ... 2 marks

13. Part of the spectrum of hydrogen is sketch below.



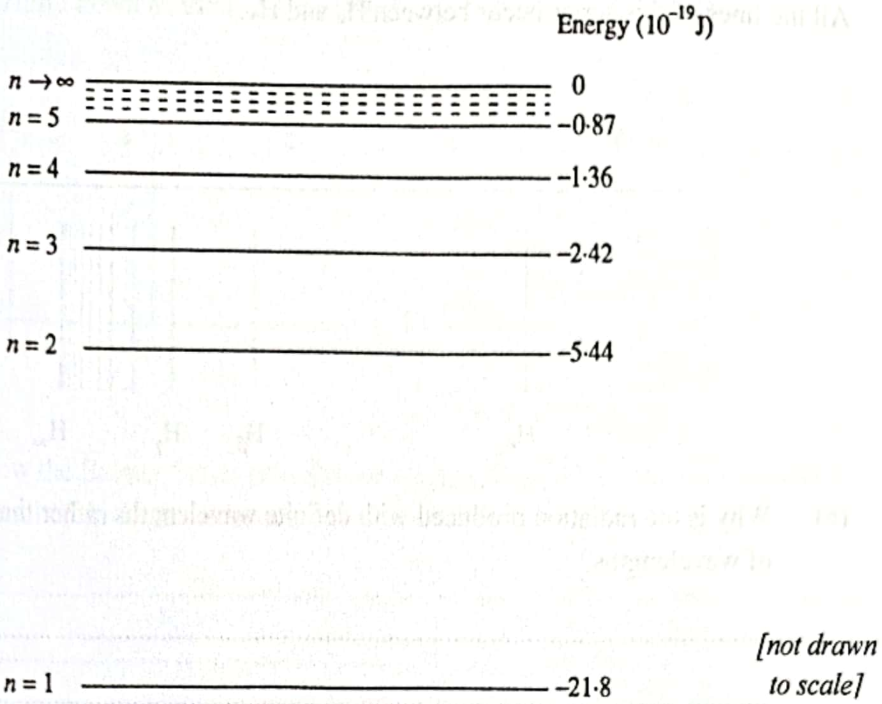
The wavelength of the lines in the series labelled *B* is given by Balmer's equation:

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

Calculate the wavelength of the line marked × in the hydrogen spectrum shown above.

1998 HSC Q32C (c)(iii) ... 2 marks

14. The diagram below represents the energy levels of a hydrogen atom.



(a) On the diagram above, indicate the transitions that represent three lines from the Balmer series emission spectrum.

(b) Suppose that a hydrogen atom in the ground state ($n = 1$) undergoes excitation by a particle whose energy is 18.0×10^{-19} J. To which energy state is the atom excited? Explain your reasoning.

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(c) Calculate the frequency and wavelength of light produced by an electron dropping from the fourth to the second energy level.

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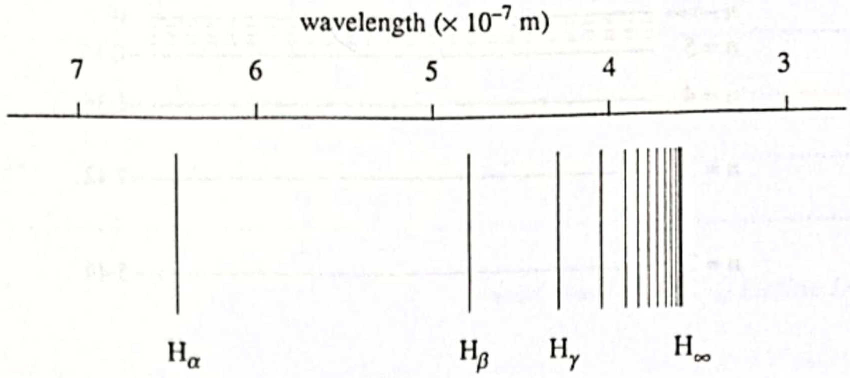
Adapted from 1992 HSC Q Elective 1A(d)
& Adapted 1996 Q32(d)(i) ... 1 + 2 + 2 = 5 marks

15. The Balmer spectral series produced by hydrogen is represented in the diagram. All the lines of this series occur between H_α and H_∞.



(a) Why is the radiation produced with definite wavelengths rather than a continuous range of wavelengths?

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(b) The wavelength, λ, of each of the above lines in the Balmer series is given by the equation:

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

What is the value of n for the H_β line? ....

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1993 HSC Q Elective 1A (c) ... 2 + 1 = 3 marks

16. The emission spectra produced by heated gases were well known when in 1885 Balmer developed an equation showing the relationship between the four lines in the visible light region of the hydrogen spectrum. His equation for this was modified by Rydberg to:

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

For the Balmer series $n_f = 2$, calculate the longest wavelength of this spectral series.

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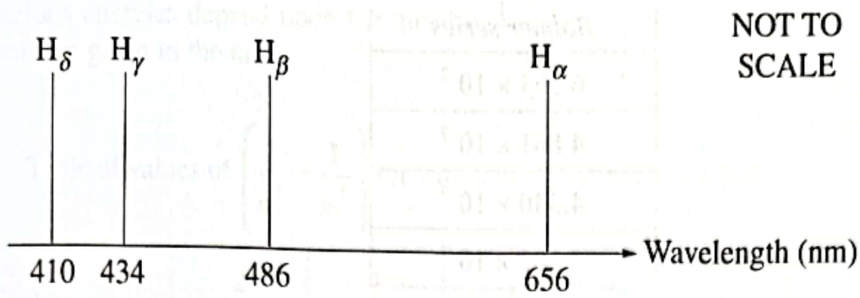
1988 HSC Q Elective 1A (c) ... 2 marks

17. The diagram represents the four spectral lines in the visible region of the hydrogen spectrum known as the Balmer Series.



(a) Explain how the Balmer Series provides strong experimental evidence in support of Bohr's model of the hydrogen atom.

(b) Calculate the wavelength of the next line in the Balmer Series.

2002 HSC Q31(c) ... 3 + 3 = 6 marks

[Part (a) is similar to: 2012 HSC Q34(a)(i)]

18. When hot hydrogen vapour is viewed through a spectrometer, four distinct bright lines of the Balmer series can be seen. The wavelengths of these lines are given in the table below.

Balmer series
6.563×10^{-7}
4.861×10^{-7}
4.340×10^{-7}
4.102×10^{-7}

(a) By referring to the Bohr model of the atom, describe how these lines are formed.

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(b) Using the Rydberg equation, show how any ONE of the above spectral lines can be predicted.

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1997 HSC Q32C(e) ... 2 + 2 = 4 marks

19. Use Balmer's equation to calculate the wavelength and energy of the photon emitted when the electrons in an excited hydrogen atom make a transition from the $n = 4$ state to the $n = 2$.

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20. The Bohr picture of the atom explains the energy of the photons emitted when an electron falls from an initial orbit n_i to find an orbit of n_f.

The various energies depend upon the quantity $\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$ and the values for this term are given in the table.

Table of values of $\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$ for values of n_i and n_f from 1 to 6

n_f \ n_i	1	2	3	4	5	6
1	–	0.7500	0.8889	0.9375	0.9600	0.9722
2	–	–	0.1389	0.1875	0.2100	0.2222
3	–	–	–	0.0486	0.0711	0.0833
4	–	–	–	–	0.0225	0.0347
5	–	–	–	–	–	0.0122
6	–	–	–	–	–	–

(a) Identify the physical reason for about one-half of the table appearing blank.

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(b) Calculate the energy of the photon emitted when an electron falls from $n_i = 4$ to $n_f = 3$.

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2006 HSC Q31(a) ... 2 + 3 = 5 marks

21. Calculate the wavelength of a photon which is emitted when an electron ... moves from energy level $n = 4$ to $n = 2$.

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2010 HSC Q36(b) ... 2 marks

22. A photon is emitted when an electron in a hydrogen atom transitions from the $n = 3$ excited state to the ground state. Calculate the wavelength of the photon.

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2014 HSC Q35(a)(i) ... 2 marks

23. Calculate the initial energy level of an electron in a hydrogen atom if it emitted 4.089×10^{-19} J on transition to the $n = 2$ level.

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2016 HSC Q34(b)(ii) ... 3 marks

21. Calculate the wavelength of a photon which is emitted when an electron in a hydrogen atom moves from energy level $n = 4$ to $n = 2$.

2010 HSC Q36(b) ... 2 marks

22. A photon is emitted when an electron in a hydrogen atom transitions from the $n = 3$ excited state to the ground state. Calculate the wavelength of the photon.

2014 HSC Q35(a)(i) ... 2 marks

23. Calculate the initial energy level of an electron in a hydrogen atom if it emitted 4.089×10^{-19} J on transition to the $n = 2$ level.

2016 HSC Q34(b)(ii) ... 3 marks

24. Outline the experimental evidence that confirmed the de Broglie hypothesis of wave-particle duality.

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2002 HSC Q31(a) ... 2 marks

25. During your Physics course, you will have investigated the spectrum of atomic hydrogen. The four coloured lines are listed in the table below.

Colour of the emission line	*Name of the emission line*	*Electron transition*
Red	H_α	$n = 3$ to $n = 2$
Green	H_β	$n = 4$ to $n = 2$
Blue	H_γ	$n = 5$ to $n = 2$
Violet	H_δ	$n = 6$ to $n = 2$

Calculate the wavelength of the H_β spectral line, and hence determine the energy of the emitted photon.

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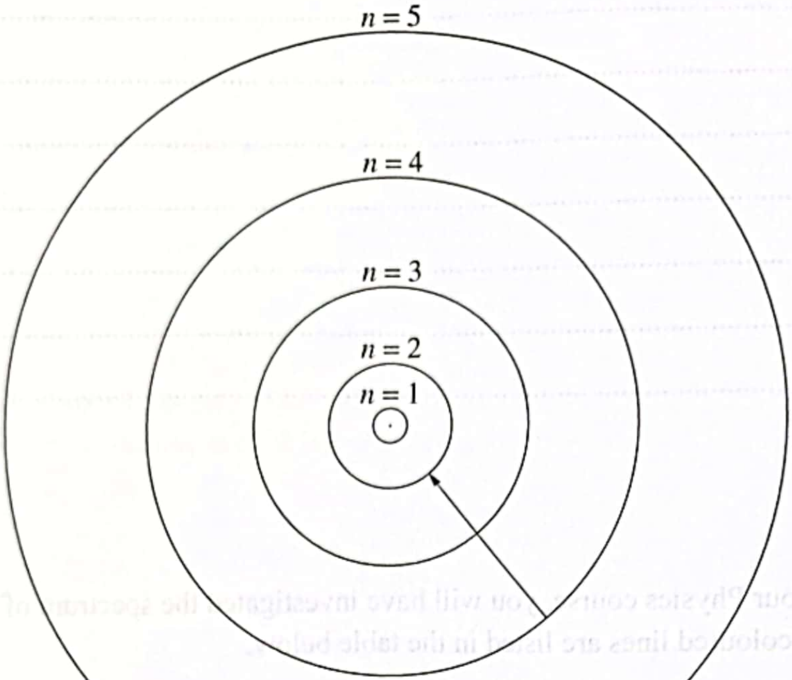
Adapted 2004 HSC Q31(d)(ii) ... 4 marks

26. The diagram below shows the first five circular Bohr orbits or 'stationary states' for the electron orbiting the nucleus of the hydrogen atom.



(a) For the electron transition shown on the diagram, calculate the wavelength of the emitted photon.

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(b) Using de Broglie's formula, calculate the wavelength of the electron in the first stationary state if its speed is 2.188×10^6 m s^{-1}.

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Adapted 2005 HSC Q31(d)(i), (ii) ... 2 + 2 = 4 marks

27. Calculate the velocity of an electron that has a wavelength of 3.33×10^{-10} m.

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2009 HSC Q31(b)(iii) ... 2 marks

28. Outline de Broglie's proposal about the properties of particles in a matter wave.

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29. (a) Which physicist proposed the concept of electron probability clouds?

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(b) Outline how the 'electron cloud' model of the atom differs from Bohr's model.

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... 1 + 2 = 3 marks

30. A cricket ball has a mass of 156 g and a velocity of 20 m s^{-1}.

Calculate its de Broglie wavelength.

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2011 HSC Q34(d) ... 2 marks

31. Why do everyday objects, such as a cricket ball, not appear to have a wavelength.

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8.3 Quantum Mechanical Nature of the Atom

Multiple choice: 1 mark each

1. B **2.** A **3.** B **4.** A **5.** D **6.** C

Explanations:

1. **B** Photon energy, $E = hf = \frac{hc}{\lambda}$ since $c = f\lambda$

$$\therefore E = \frac{6.626 \times 10^{-34}\ \text{J s} \times 3.00 \times 10^{8}\ \text{m s}^{-1}}{3.5 \times 10^{-2}\ \text{m}}$$

$$= 5.68 \times 10^{-24}\ \text{J} \ldots \text{as in (B).}$$

2. **A** $E = hf$ and $c = f\lambda$

$$\ldots \text{so } E = h\frac{c}{\lambda} = 6.626 \times 10^{-34} \times \frac{3.00 \times 10^{8}}{580 \times 10^{-9}} = 3.427 \times 10^{-19}\ \text{J} \approx 3.43 \times 10^{-19}\ \text{J}$$

So (A) is the answer.

3. **B** $E = hf$ and $c = f\lambda \quad \therefore E = h\frac{c}{\lambda}$

$$3.5 \times 1.602 \times 10^{-19} = 6.626 \times 10^{-34} \times \frac{3.00 \times 10^{8}}{\lambda}$$

$$\therefore \lambda = 6.626 \times 10^{-34} \times \frac{3.00 \times 10^{8}}{3.5 \times 1.602 \times 10^{-19}} = 3.5 \times 10^{-7}\ \text{m} \quad \ldots \text{as in (B).}$$

4. **A** Davisson and Germer designed and built a vacuum apparatus to observe the intensity of electrons scattered from a polished nickel target at varying angles of incidence. They discovered that the intensity peaked regularly at particular angles of incidence and reduced to a minimum at angles in-between. This interference pattern showed the wave nature of the electrons (and hence the existence of matter waves). The wavelength of the electron waves, calculated from the observed interference patterns, agreed exactly with the wavelength calculated using $\lambda = \frac{h}{mv}$. This confirmed de Broglie's equation and verified his prediction. So (A) is the answer and (B) is incorrect.

Schrödinger's wave equations applied to atomic electrons, rather than electron beams, so (C) is incorrect. The Rutherford-Bohr atomic model and Balmer series were unrelated to wave properties of electrons, so (D) is incorrect.

5. **D** $c = f\lambda$

$$\therefore f = \frac{c}{\lambda} = \frac{3.00\times10^{8}}{550\times10^{-9}}$$

$= 5.45 \times 10^{14}$ Hz ... so (D) is the answer.

6. **C** Schrödinger's model had electrons moving around the nucleus as an electron cloud due to their wave nature, where the location of an electron at any point in time is based on probability. The cloud density was greatest where you would be most likely to find an electron. Whereas Bohr had electrons in well-defined circular paths around the nucleus. So (C) is the answer and (B) is incorrect. Schrödinger's thought experiment was an approach to explain quantum mechanical systems and was not a model of the atom. So (A) is incorrect. Schrödinger's model applies to atoms other than the hydrogen atom. So (D) is incorrect.

[Note: Interestingly, Bohr subsequently developed his complementarity principle (in 1927). He stated that some mutually exclusive views of nature could *both* be true, but not at the same time, i.e. light could be a wave in one experiment, but particles in another experiment. However, light could not show both wave and particle properties at the same time.]

Short-answer questions

7. (a) The Rutherford Model described the atom as having a tiny, central positive nucleus. The rest of the atom was mostly empty space, occupied by negatively charged electrons that orbited at a considerable distance from the nucleus.

(b) LIMITATION: Rutherford's model could not explain why negative electrons stayed away from the positive nucleus without collapsing into it.

Bohr overcame this limitation by introducing quantum theory to Rutherford's model. He proposed that electrons could only occupy orbits of specific radius and fixed energy levels. Bohr explained that electrons are in a stable 'stationary state' and do not radiate energy while in a particular orbit.

OR LIMITATION: Rutherford's model could not explain the hydrogen emission lines.

Bohr overcame this limitation by placing electrons into quantised energy shells. When electrons moved from a higher to a lower shell, they released energy that corresponded to the frequencies in the emission spectra.

8. *Any TWO of the following:*

- It only explains spectra of 'hydrogen-like' species, e.g. H, He^{+} (not multi-electron atoms).
- It does not account for the fine, closely spaced spectral lines (hyperfine spectral lines).
- It provides no explanation for the different intensities of spectral lines.
- It does not explain magnetic splitting of spectral lines (the Zeeman effect).
- It does not explain molecular bonding.

9. It did not explain the stability of an atom. In classical physics, according to the law of conservation of energy, an electron orbiting a nucleus should emit energy due to its motion. That energy loss would in turn cause the electron to spiral into the nucleus and the atom would collapse.

10. (a) They would have examined it using a spectroscope. They would have seen how it consisted of four bright lines of visible light – a red line, a green line, a blue line and a violet line on a dark background.

(b) In Bohr's model of the atom, an electron can become excited to a higher energy state and then, as it moves back to a lower level, it emits an exact amount of energy in the form of photon of light. The emission wavelengths of hydrogen seen in this investigation would agree with those calculated using Bohr's model of the atom.

11. Bohr proposed that electrons were in a specific energy level in a stable orbit around the nucleus. An excited electron in a hydrogen atom can fall from a higher energy level to a lower energy level and in doing so, it emits a photon of specific energy (with $E = hf$). This results in the emission spectrum of hydrogen.

12. Bohr accounted for the stability in atomic structure by proposing that electrons existed in stable stationary states in which an electron could orbit the nucleus without emitting electromagnetic radiation, as required by Rutherford's model.

13. From $\frac{1}{\lambda} = R\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right]$

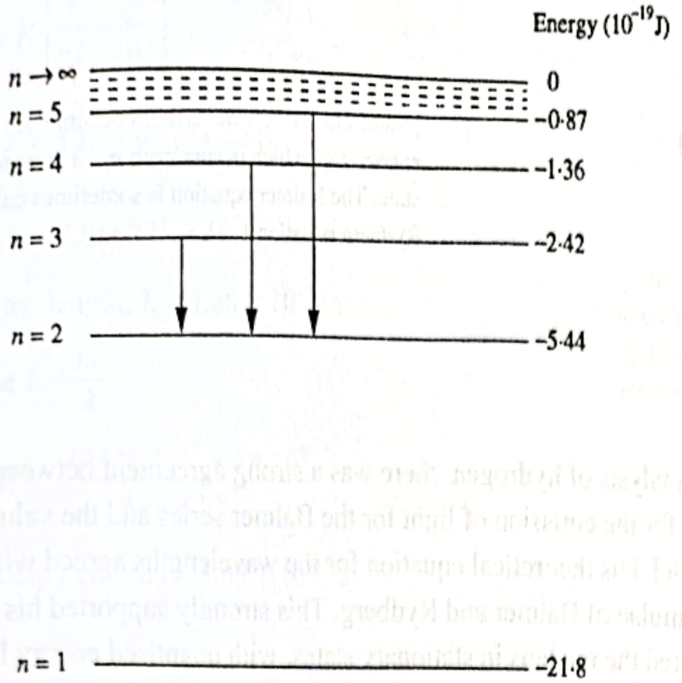
$\frac{1}{\lambda} = 1.097 \times 10^7 \left(\frac{1}{2^2} - \frac{1}{3^2}\right)$

$= 1.524 \times 10^{-6}$

$\therefore \lambda$ for line at **×** $= 6.56 \times 10^{-7}$ m

[Note: The 4 visible lines in the Balmer series correspond to $n_f = 2$ and $n_i = 3, 4, 5$ or 6. Since **×** is pointing to the longest wavelength in the *B* series, $n_i^2 = 3$ … as the longest wavelength occurs with the smallest energy gap, i.e. from $n = 3$ state to $n = 2$ state.]

14. (a)



[Note: The Balmer series shows transitions back to $n = 2$.]

(b) *Using the data in the diagram:*

To transition from $n = 1$ to $n = 2$, energy required $= -21.8 - (-5.44) = 16.36 \times 10^{-19}$ J and from $n = 1$ to $n = 3$, energy required $= -21.8 - (-2.42) = 19.38 \times 10^{-19}$ J

$\therefore$ if the energy available is 18.0×10^{-19} J, it does not correspond to either of these transitions. So, the energy is not absorbed. Hence the atom remains in the ground state.

(c) $\Delta E_{(4\text{ to }2)} = (-1.36 \times 10^{-19}) - (-5.43 \times 10^{-19})$

$= 4.07 \times 10^{-19}$ J

Calculating f: $\Delta E = hf$

So $4.07 \times 10^{-19} = 6.626 \times 10^{-34} \times f$

$\therefore$ frequency, $f = \dfrac{4.07 \times 10^{-19}}{6.626 \times 10^{-34}} = 6.14 \times 10^{14}$ Hz

Calculating λ: $\Delta E = \dfrac{hc}{\lambda}$

So $4.07 \times 10^{-19} = \dfrac{6.626 \times 10^{-34} \times 3.00 \times 10^{8}}{\lambda}$

$\therefore$ wavelength, $\lambda = \dfrac{6.626 \times 10^{-34} \times 3.00 \times 10^{8}}{4.07 \times 10^{-19}} = 4.88 \times 10^{-7}$ m

15. (a) Each line is produced as an electron drops back down to one of the lower energy states. Each spectral line has a distinct wavelength that is equivalent to the energy difference between the initial and final energy state of the electron.

(b) The ground state is $n = 2$, so H_α would be $n = 3$ and H_β is $n = 4$.

16. From $\frac{1}{\lambda} = R\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right]$

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

$= 1.524 \times 10^{-6}$

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

[Note: The longest wavelength occurs with the smallest energy gap, which occurs from $n = 3$ state to $n = 2$ state. The Balmer equation is sometimes called the Rydberg equation.]

17. (a) When Bohr did his analysis of hydrogen, there was a strong agreement between the experimental values for the emission of light for the Balmer series and the values predicted by his model. His theoretical equation for the wavelengths agreed with the earlier empirical formulae of Balmer and Rydberg. This strongly supported his model in which electrons orbited the nucleus in stationary states, with quantised energy levels and how light of a specific frequency is emitted or absorbed when an electron makes a transition between one energy level and another. The frequency of each spectral line explained the transition of electrons between a pair of energy levels in the H atom.

(b) $\frac{1}{\lambda} = R\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$ where $n_f = 2$, and $n_i =$ 3, 4, 5, 6, 7, etc

For the next line in the series, $n_i = 7$

$\frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{7^2}\right) = R(0.25 - 0.020408) = 1.097 \times 10^7(0.2296) = 2.519 \times 10^6$

$\therefore$ wavelength of next line, $\lambda = 3.97 \times 10^{-7}$ m = 397 nm

18. (a) Bohr's model of the hydrogen atom proposed that its electron orbited the central nucleus in a stable stationary state that was the lowest energy state. When the atom is excited, the electron can move to a higher energy level and when it drops back down to a lower state, energy is released. This corresponds to the difference in energy between the two states and gives rise to specific spectral lines. When the electrons drop from a higher energy state to the $n = 2$ state, the Balmer series of spectral lines is produced.

(b) $\frac{1}{\lambda} = R\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right]$ *[Note: The Rydberg equation = the Balmer equation]*

In the Balmer series, $n_f = 2$.

For the spectral line produced by a drop from $n_i = 5$: $\frac{1}{\lambda} = 1.097 \times 10^7 \left(\frac{1}{2^2} - \frac{1}{5^2}\right)$

$= 2.3037 \times 10^6$

$\therefore \lambda$ for this spectral line $= 4.34 \times 10^{-7}$ m

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

So $\frac{1}{\lambda} = 1.097 \times 10^7 \left(\frac{1}{2^2} - \frac{1}{4^2}\right)$

$= 2.056875 \times 10^6$

$\therefore$ wavelength, $\lambda = 4.86 \times 10^{-7}$ m

[Note: This photon will have a shorter wavelength than going from level 3 to level 2, and would be in the UV region of the electromagnetic spectrum.]

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

So $E = \frac{6.626 \times 10^{-34} \times 3 \times 10^8}{4.86 \times 10^{-7}}$

$\therefore$ energy emitted $= 4.09 \times 10^{-19}$ J

20. (a) Photons are only emitted when an electron loses energy as it moves from a higher to a lower energy orbit. The blank spaces to the immediate left of the first number in each row represent no change in energy level, i.e. 6→6, 5→5, etc. The remaining blank spaces represent orbit changes from a lower to a higher energy level that would require the absorption of a photon, not the emission of a photon.

(b) Using Balmer's equation: $\frac{1}{\lambda} = R\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right]$

For the electron transition, $n_i = 4$ to $n_f = 3$ (from data table):

$\frac{1}{\lambda} = R_H \times 0.0486 = 1.097 \times 10^7 \times 0.0486 = 5.33142 \times 10^5 \text{ m}^{-1}$

$E = \frac{hc}{\lambda} = hc \times \frac{1}{\lambda} = 6.626 \times 10^{-34} \times 3.00 \times 10^8 \times 5.33142 \times 10^5 = 1.05978 \times 10^{-19}$ J

$\therefore$ Energy of photon, $E = 1.06 \times 10^{-19}$ J

21. $\frac{1}{\lambda} = R\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) = 1.097 \times 10^7 \times \left(\frac{1}{2^2} - \frac{1}{4^2}\right) = 1.097 \times 10^7 \times \left(\frac{1}{4} - \frac{1}{16}\right)$

So $\frac{1}{\lambda} = 1.097 \times 10^7 \times \frac{3}{16} = 2.056875 \times 10^6$

$\therefore \lambda = (2.056875 \times 10^6)^{-1} = 4.8617 \times 10^{-7}$ m

[Note: This is the H_β line in the Balmer series of the hydrogen spectrum.]

22. (a) $\frac{1}{\lambda} = R\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) = 1.097\times10^7 \times\left(\frac{1}{1^2} - \frac{1}{3^2}\right) = 1.097\times10^7\times(1-0.1111)$

$= 1.097 \times 10^7 \times 0.8889 = 9.751 \times 10^6 \text{ m}^{-1}$

$\therefore$ wavelength of photon, $\lambda = 1.0255 \times 10^{-7}$ m = 102.6 nm

23. Since $E = hf$ and $E = \frac{hc}{\lambda}$ and Balmer's equation is: $\frac{1}{\lambda} = R\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$

$\therefore \frac{E}{hc} = R\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$ and so: $\frac{1}{n_i^2} = \frac{1}{n_f^2} - \frac{E}{hcR}$

$\frac{1}{n_i^2} = \frac{1}{2^2} - \frac{4.089\times10^{-19}}{6.626\times10^{-34}\times3.00\times10^8\times1.097\times10^7}$ so $\frac{1}{n_i^2} = 0.06248$

$\therefore n_i = \sqrt{\frac{1}{0.06248}} = 4$ i.e. initial energy level of an electron = 4

24. Davisson and Germer accelerated electrons towards a nickel crystal. They observed that the electrons were being diffracted as they scattered off the solid nickel crystal surface. This produced an interference pattern. As diffraction is a property of waves and not particles, they had confirmed that electrons had the wave nature proposed by de Broglie. The wavelength of the electrons, calculated from the observed interference patterns using $\lambda = \frac{h}{mv}$, agreed exactly with de Broglie's predicted value.

25. Using Balmer's equation: $\frac{1}{\lambda} = R_H\left(\frac{1}{n_f^{\,2}} - \frac{1}{n_i^{\,2}}\right)$

For the green H_β spectral line: $n_i = 4$ and $n_f = 2$ *[from data table]*

$\frac{1}{\lambda} = 1.097\times10^7\left(\frac{1}{2^2} - \frac{1}{4^2}\right) = 1.097\times10^7\times\left(\frac{1}{4} - \frac{1}{16}\right) = 2.056875\times10^6 \text{ m}^{-1}$

$\lambda = (2.056875\times10^6)^{-1} = 4.8617\times10^{-7}$ m

$\therefore$ wavelength of H_β spectral line, $\lambda = 4.862 \times 10^{-7}$ m

$E = hf = \frac{hc}{\lambda} = \frac{6.626 \times 10^{-34} \times 3.00 \times 10^8}{4.8617 \times 10^{-7}} = 4.0887\times10^{-19}$ J

$\therefore$ energy of emitted photon, $E = 4.089 \times 10^{-19}$ J

26. (a) Using Balmer's equation: $\frac{1}{\lambda} = R\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right]$

For the line shown, $n_f = 2$ and $n_i = 4$ (from diagram)

$$\frac{1}{\lambda} = 1.097 \times 10^7\left(\frac{1}{2^2} - \frac{1}{4^2}\right) = 1.097 \times 10^7\left(\frac{1}{4} - \frac{1}{16}\right) = 2.056875 \times 10^6\ \text{m}^{-1}$$

$$\lambda = (2.056875 \times 10^6)^{-1} = 4.8617 \times 10^{-7}\ \text{m}$$

$\therefore$ wavelength of emitted photon line, $\lambda = 4.862 \times 10^{-7}$ m

(b) de Broglie's formula is: $\lambda = \frac{h}{mv}$

If an electron is travelling at 2.188×10^6 m s^{-1}:

$\therefore$ wavelength of electron, $\lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34}}{9.109 \times 10^{-31} \times 2.188 \times 10^6} = 3.325 \times 10^{-10}$ m

27. $\lambda = \frac{h}{mv}$ So $v = \frac{h}{m\lambda}$

$$\therefore v = \frac{h}{m\lambda} = \frac{6.626 \times 10^{-34}}{9.109 \times 10^{-31} \times 3.33 \times 10^{-10}} = 2.184 \times 10^6\ \text{m s}^{-1}$$

28. de Broglie suggested that if light can behave as both a wave and as a photon, particles in a matter wave could also have dual character. So, he proposed that all particles had wave properties as well as particle properties and that the wavelength is inversely proportional to the momentum of the particle, i.e. $\lambda = \frac{h}{mv}$.

29. (a) Schrödinger [OR Erwin Schrödinger]

(b) It has the electrons moving around the nucleus as an electron cloud, where the position of an electron at any point in time is based on probability. The cloud density was greatest where you would be most likely to find an electron. This is because electrons are considered to have a wave nature and not to be a particle. Whereas the electrons in the Bohr model have a clearly defined orbit, in which the electrons surround the nucleus.

30. $\lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34}}{0.156 \times 20} = 2.12 \times 10^{-34}$ m

31. Everyday objects can have momentum and so possess wave characteristics. However, the wave nature of larger objects is far too small to be observed or measured due to their extremely short wavelengths. Hence their wave nature is usually disregarded.