

3.4 THE BOHR ATOMIC THEORY

PRACTICE

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- | | | |
|---|---|--|
| 1. (a) | (b) | (c) |
| $3 e^{-}$
$2 e^{-}$
$5 p^{+}$
B
boron atom | $3 e^{-}$
$8 e^{-}$
$2 e^{-}$
$13 p^{+}$
Al
aluminum atom | $2 e^{-}$
$2 p^{+}$
He
helium atom |

SECTION 3.4 QUESTIONS

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Understanding Concepts

- The main achievement of the Rutherford model was the advancement of atomic theory to include the nucleus. The main problem was explaining the region occupied by the electrons. According to existing theory, electrons should spiral into the nucleus as they lose energy by emitting electromagnetic radiation.
- Bohr's solution to the problem with the Rutherford model was to assume that classical ideas of energy did not hold inside the atom, and that electron energies were quantized in special energy states called stationary states.
- The important new idea used by Bohr was the quantum theory of light, together with the experimental evidence of line spectra.
- (a) Emission spectra consist of light emitted by a sample of a substance. Absorption spectra consist of the missing frequencies (colours) of light after the source light passes through a sample of a substance.
 (b) Bohr explained emission spectra as light emitted by an atom when its electrons drop from higher to lower energy states. The dark lines in absorption spectra correspond to the light absorbed by electrons jumping from lower to higher energy states.
- Niels Bohr and Ernest Rutherford both made major contributions, at approximately the same time, to our understanding of the structure of the atom. Rutherford's work described the nucleus, and Bohr's work described the extranuclear region of the atom.
- | | | |
|--|--|---|
| (a) | (b) | (c) |
| $7 e^{-}$
$2 e^{-}$
$9 p^{+}$
F
fluorine atom | $8 e^{-}$
$2 e^{-}$
$10 p^{+}$
Ne
neon atom | $1 e^{-}$
$8 e^{-}$
$2 e^{-}$
$11 p^{+}$
Na
sodium atom |

- Atomic numbers give the number of protons, and for neutral atoms, the number of electrons as well. The period number shows how many levels of electron energy there are, and the last digit of the group number shows how many electrons are in the valence (highest) energy level.
- Bohr's theory was considered a success because it explained the known emission spectral lines for hydrogen, and predicted successfully some lines in the infrared light spectrum. Bohr's theory also provided a better understanding of the arrangement of elements in the periodic table.
- One significant problem with the Bohr theory was that it did not predict correctly the spectral lines for atoms with more electrons than hydrogen. (It also could not explain some simple observations such as some emission lines were brighter than others.)

Applying Inquiry Skills

- Element 118, based on the periodic law and the Bohr theory of the atom, should be in group 18, making it a noble gas (unreactive chemically) with a full valence electron level (18 electrons in the seventh level), with a very high density for a gas, around 14 ± 3 g/L, and melting and boiling points of around $-20 \pm 20^{\circ}\text{C}$.

Making Connections

- (Some of the content that students could recognize include references to spectra of elements, arrangement of electrons in shells (2, 8, 18, 32), stationary states, and transitions between states. Much of the discussion about this content, and

other topics, will be somewhat obscure. At best, probably about three-quarters of the total discussion is beyond students' current understanding. This is a good illustration of how content, even of an apparently "simple" theory like Bohr's, is considerably simplified and restricted at the high school level.)

12. To determine if a sample of table salt, $\text{NaCl}_{(s)}$, contains some potassium chloride, a flame test can be done and the characteristic violet flame will indicate if potassium is present. (The violet colour of potassium is normally overwhelmed by the yellow of sodium; but the colours can be separated by using a spectroscope to view the light, or by viewing the light through deep-blue cobalt glass which absorbs the yellow sodium wavelengths but not the violet potassium wavelengths.)

Extensions

13. (a) $\frac{1}{\lambda} = R_H \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$

$$n_f = 2 \quad n_i = 4$$

$$R_H = 1.10 \times 10^7 / \text{m}$$

$$\frac{1}{\lambda} = \frac{1.10 \times 10^7}{1 \text{ m}} \left(\frac{1}{2^2} - \frac{1}{4^2} \right)$$

$$\lambda = \frac{1 \text{ m}}{2.06 \times 10^6}$$

$$\lambda = 4.85 \times 10^{-7} \text{ m} = 485 \text{ nm}$$

The wavelength of light emitted is 485 nm.

(b) $\lambda = \frac{c}{f}$

$$c = 3.00 \times 10^8 \text{ m/s}$$

$$f = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \text{ m/s}}{4.85 \times 10^{-7} \text{ m}}$$

$$f = \frac{6.18 \times 10^{14}}{1 \text{ s}} = 6.18 \times 10^{14} \text{ s}^{-1}$$

The frequency of the light emitted is $6.18 \times 10^{14} \text{ s}^{-1}$.

(c) $E = hf$

$$h = 6.63 \times 10^{-34} \text{ J/Hz}$$

$$E = \frac{6.63 \times 10^{-34} \text{ J}}{1 \text{ Hz}} \times 6.18 \times 10^{14} \text{ Hz}$$

$$E = 4.1 \times 10^{-19} \text{ J}$$

The electron's energy difference between the second and fourth levels is $4.1 \times 10^{-19} \text{ J}$.

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

$$R_H = 1.10 \times 10^7 / \text{m}$$

$$n_f = 2 \quad n_i = 3$$

$$\frac{1}{\lambda} = \frac{1.10 \times 10^7}{1 \text{ m}} \left(\frac{1}{2^2} - \frac{1}{3^2} \right)$$

$$\lambda = 6.54 \times 10^{-7} \text{ m} = 654 \text{ nm}$$

The wavelength of light emitted is 654 nm.

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

$$c = 3.00 \times 10^8 \text{ m/s}$$

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

$$= \frac{3.00 \times 10^8 \text{ m/s}}{6.54 \times 10^{-7} \text{ m}}$$

$$f = \frac{4.59 \times 10^{14}}{1 \text{ s}} = 4.59 \times 10^{14} \text{ Hz}$$

The frequency of the light emitted is $4.59 \times 10^{14} \text{ Hz}$.

$$E = hf$$

$$h = 6.63 \times 10^{-34} \text{ J/Hz}$$

$$E = \frac{6.63 \times 10^{-34} \text{ J}}{1 \text{ Hz}} \times 4.59 \times 10^{14} \text{ Hz}$$

$$E = 3.0 \times 10^{-19} \text{ J}$$

The electron's energy difference between the second and third levels is $3.0 \times 10^{-19} \text{ J}$.

(e)

$$\begin{array}{lcl} \text{----} & n = 4 & \left. \vphantom{\begin{array}{l} n=4 \\ n=3 \\ n=2 \end{array}} \right\} \Delta E_{4 \rightarrow 2} = 4.1 \times 10^{-19} \text{ J} \\ \text{----} & n = 3 & \left. \vphantom{\begin{array}{l} n=4 \\ n=3 \\ n=2 \end{array}} \right\} \\ \text{----} & n = 2 & \left. \vphantom{\begin{array}{l} n=4 \\ n=3 \\ n=2 \end{array}} \right\} \Delta E_{3 \rightarrow 2} = 3.0 \times 10^{-19} \text{ J} \\ 1 \text{ e-} & n = 1 & \\ 1 \text{ p}^+ & & \\ \mathbf{H} & & \end{array}$$

hydrogen atom

$$(4.1 \times 10^{-19} \text{ J}) - (3.0 \times 10^{-19} \text{ J}) = 1.1 \times 10^{-19} \text{ J}$$

The energy difference between hydrogen atom electron energy levels $n = 4$ and $n = 3$ will be $1.1 \times 10^{-19} \text{ J}$.

14. The most likely assumption would seem to be that theories would be advanced that would try to describe electron arrangements and energies for atoms more complex than hydrogen.

3.5 QUANTUM NUMBERS

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Understanding Concepts

- Bohr and Sommerfeld both used observations of line spectra.
- Bohr proposed circular electron orbits for hydrogen, while Sommerfeld proposed several elliptical orbits.