CHEM1101 Critical Thinking Questions 3: The Energy Levels Of Electrons

Model 1: Electron Energy

For an atom with one electron orbiting around a nucleus with charge \( Z \), the energy of the electron is given by the equation below:

\[
E_n = -(2.18 \times 10^{-18} \text{ J}) \frac{Z^2}{n^2}
\]

where \( n = 1, 2, 3, 4 \ldots \). The different values of \( n \) correspond to the allowed energies that the electron can have. These energies are called “energy levels”. The lowest energy level has \( n = 1 \) and is called the “ground state”. All other energy levels are called “excited states”. The average distance of an electron from a nucleus is also controlled by the value of \( n \):

\[
r_{\text{average}} = (0.529 \times 10^{-10} \text{ m}) \frac{n^2}{Z}
\]

Critical thinking questions

1. The hydrogen atom has atomic number \( Z = 1 \). Using the equations above for the energy and average distance of an electron from a nucleus \( (r_{\text{average}}) \), complete the table below for hydrogen.

<table>
<thead>
<tr>
<th>( n )</th>
<th>( E_n ) (J)</th>
<th>( r_{\text{average}} ) (m)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>(-218 \times 10^{-20})</td>
<td>(0.529 \times 10^{-10})</td>
</tr>
<tr>
<td>2</td>
<td>(-54.5 \times 10^{-20})</td>
<td>(2.12 \times 10^{-10})</td>
</tr>
<tr>
<td>3</td>
<td>(-24.2 \times 10^{-20})</td>
<td></td>
</tr>
<tr>
<td>4</td>
<td></td>
<td></td>
</tr>
<tr>
<td>5</td>
<td></td>
<td></td>
</tr>
<tr>
<td>6</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

2. The horizontal lines on the graph on page 3 shows the energy levels for \( n = 1, 2 \) and 3. Using the values you calculated for \( E_n \), add the energy levels for \( n = 4 – 6 \).

3. Describe in words to your neighbour what happens to the energy levels and the average distance of the electron from the nucleus as \( n \) increases. When you have agreed on this, write down your description in a grammatically correct sentence below.

4. As \( n \) gets larger and larger (approaches infinity) what happens to the energy and an average distance from a nucleus?

5. How would these energy levels change for other hydrogen-like ions, e.g. \( \text{Li}^{2+} \)?
Model 2: Atomic Spectroscopy

The electron in a H atom will be in the $n = 1$ level (the “ground state”). However, if energy is provided, an excited H atom is formed in which the electron is in a level with $n > 1$. The electron in this “excited” atom quickly moves (“relaxes”) to a lower level and the excess energy is lost (“emitted”) as electromagnetic radiation.

For example, if the excited atom is formed with its electron in the $n = 3$ level, the electron can fall into the $n = 2$ or into the $n = 1$ level. This is shown by the dotted vertical lines on the graph. The energy lost is the difference between the two energy levels involved: it is equal to the length of the dotted line.

If the electron moves from (a) $n = 3$ to $n = 1$, the energy emitted is equal to $194 \times 10^{-20}$ J and (b) $n = 3$ to $n = 2$, the energy emitted is equal to $30 \times 10^{-20}$ J. These energies can be measured on the graph or calculated using the values in the table on the previous page.

The energy of the emitted light for jumps (a) and (b) is shown by the horizontal lines on the graph of the right hand side overleaf.

Critical thinking questions

6. If the excited atom is formed with the electron in the $n = 4$ level, add dotted lines to the graph on page 3, showing how the electron can relax.

7. For each of these jumps, work out the energy of the emitted light and mark it using a horizontal line on the right hand graph on page 3.

8. Repeat questions 6 and 7 for an excited atom formed with the electron in the $n = 5$ level.

9. The right hand graph represents the energy of the emitted radiation for the hydrogen atom. You have drawn its “atomic spectrum”.

   Light with energy less than $28 \times 10^{-20}$ J is the infrared region. Light with energy greater than $50 \times 10^{-20}$ J is ultraviolet. Visible light lies between these values. Mark these regions on your spectrum.

10. (a) What wavelengths correspond to the upper and lower limits of visible light given above?

    (b) What colours correspond to these limits (see table below)?

11. What do the lines in the visible part of the spectrum have in common?

12. What do the lines in the ultraviolet part of the spectrum have in common?

<table>
<thead>
<tr>
<th>Colour</th>
<th>red</th>
<th>orange</th>
<th>yellow</th>
<th>green</th>
<th>blue</th>
<th>indigo</th>
<th>violet</th>
</tr>
</thead>
<tbody>
<tr>
<td>Wavelength (nm)</td>
<td>660</td>
<td>610</td>
<td>580</td>
<td>540</td>
<td>470</td>
<td>440</td>
<td>410</td>
</tr>
</tbody>
</table>
The energy levels of the H atom

Energy levels of the H atom

- $n = 1$
- $n = 2$
- $n = 3$

Atomic Spectrum of the H atom

(a) $n = 3 \rightarrow n = 1$

(b) $n = 3 \rightarrow n = 2$
**Model 3: Light and Waves**

The picture below shows a light wave. The wavelength is the distance between peaks (or the distance between troughs). The amplitude is the height of the wave counted from 0 (= height of a peak or depth trough). We cannot see these waves. Instead, our eyes detect the *intensity* of the light which is given by *square* of the wave.

Squaring means multiplying the wave at each point by itself, remembering that positive × positive and negative × negative are both positive.

**Critical thinking questions**

13. On the diagram, indicate the wavelength \((\lambda)\) and the amplitude \((A)\) of the wave.

14. Put an asterisk (‘*’”) to mark the positions where the wave changes sign. These are ‘nodes’.

15. Peaks are where the wave is positive. Troughs are where the wave is negative. Labels these with ‘+’ and ‘−’ signs respectively. *Lightly* shade the ‘−’ areas.

16. On top of the picture, draw a sketch of the *intensity* of the light.