Level 3 Physics: Atoms – Spectra - Answers

In 2013, **AS 91525** replaced **AS 90522.** Prior to 2013, this was an external standard - AS90522 Atoms, Photons and Nuclei.

It is likely to be assessed using an internal test from 2013 onwards (although teachers can select from a range of assessment techniques). There were only minor changes to this existing material in the standard when it became AS91525 but also a number of additions including Relativity and some material on fundamental particles. The old external examinations may be useful revision for an internal test.

However, the mess that is NCEA Assessment Schedules….

For 90522 there was an Evidence column with the correct answer and Achieved, Merit and Excellence columns explaining the required level of performance to get that grade. Each part of the question (row in the Assessment Schedule) contributed a single grade in either Criteria 1 (Explain stuff) or Criteria 2 (Solve stuff). From 2003 to 2008, the NCEA shaded columns that were not relevant to that question (Sorry haven't had time to do 2004 yet).

Question	Evidence	Achievement	Merit	Excellence
2012(2) (d)(i)	The sun emits a continuous spectrum of visible light. The dark lines are caused by absorption of specific photons. The gases in the atmosphere can only absorb photons which have energies that exactly match the difference in the energy levels of the electrons in the gases. Since the energy of a photon depends on its frequency (wavelength) and therefore its colour, only those specific colour photons which match will be absorbed, leaving a sharp dark line in the full spectrum.	[1] Light / colours / wavelengths / frequencies / energy are absorbed by gases / electrons / atoms OR Only light of **specific / certain** colour / wavelength / frequency / energy is absorbed OR Light / wavelengths / frequencies / colours / energies allow electrons to move to higher energy levels / gain energy Accept "used up" / "removed" instead of "absorbed	[1] **Photons** of specific / certain wavelength / frequency / colour / energy are absorbed by electrons which gain energy / move / jump to a higher energy level	

(ii)	The quantised energy levels for each element are fixed and distinct. So the difference in energy between the levels is also fixed and distinctive for a given element. Electrons can move up (absorb photons) as well as move down (emit photons) these distinct energy levels and so they will emit photons at the same colours (frequencies / wavelengths / energies) as they will absorb them. So since the dark lines appear at the precisely the colours (wavelengths / frequencies / energies) that occur in 'sodium light', the absorption at these matching colours indicates sodium is present in the sun's atmosphere.	[1] Photon / wavelength / frequency linked directly to sodium / unique spectrum of sodium. OR Emission and absorption lines match.	[1] **Photon absorbed / emitted** or **energy levels** linked to sodium. AND Emission and absorption lines match.	
For all (c) - d(ii)	In questions (c), d(i) or d(ii) students need to show that: • Electrons exit in discrete / quantised energy levels around the atom of a given gas. • Electrons must lose or gain energy to move up or down energy levels. • Electrons absorb or emit photons in order to gain or lose energy and so move up and down energy levels. • The energy of the absorbed or emitted photon must match exactly the difference in energy level of the electron.	Electrons exist in energy levels. OR Electrons lose or gain energy by emitting or absorbing photons / light / frequencies / wavelengths / colours.	Electrons exist in **discrete / quantised / specific / certain** energy levels. OR The difference in energy levels of the electron corresponds to the **photon** energy released or absorbed.	The difference in the **discrete / quantised / specific / certain** energy levels corresponds to the **photon** energy emitted or absorbed.

2011(2) (a)	A photon is a packet/particle/quantum of (electromagnetic) energy/light It can be produced when an electron in a higher energy level drops to a lower energy level.	[1]Photon is a packet/particle of light / energy. (Do **not** allow "ray of energy", "particle that holds energy".) OR Photon release linked to energy levels.	[1]Photon is a packet / particle of light / energy. (Do **not** allow "ray of energy", "particle that holds energy".) AND Photons are released when an **electron** drops to a lower energy level.	
2010(1) (a)(i)	Circles are **energy levels** (of electrons) or *electron* shells or orbits. Not atoms. Not just shells Lines represent **electron transitions** between energy levels, or spectral lines (not electromagnetic spectrum), wavelengths emitted (or absorbed)	[1]Correct identification of energy levels and electron transitions.		
(ii)	Emission spectra explained by: a transition from higher to a lower energy level causing the emission of a photon of a discrete energy level. Only certain energies are involved, so the photons emitted have certain energies / wavelengths – hence the lines in the spectrum. Absorption spectra explained by: a transition from lower to a higher energy levels caused by the absorption of only specific photons with the 'right' energies – hence the dark lines in the spectrum. The wavelength corresponds to the CHANGE in energy level.	[1]Describes emission or absorption spectra in terms of electron transitions.	[1]Explains emission OR absorption spectra in terms of electron transitions of **discrete energies.**	[1]must include that wavelength (energy released or absorbed)is directly proportional to **Change in Energy.**

(b)	$\frac{1}{\lambda} = R\left(\frac{1}{S^2} - \frac{1}{L^2}\right)$ $\frac{1}{\lambda} = 1.097 \times 10^7 \left(\frac{1}{2^2} - \frac{1}{5^2}\right)$ $= 2.304 \times 10^6$ $\lambda = 4.341 \times 10^{-7}$ m	[2]Correct method.		
(c)	$E_n = \frac{hcR}{n^2}$ $E_1 = -\frac{6.626 \times 10^{-34} \times 3.00 \times 10^8 \times 1.097 \times 10^7}{1^2}$ $E_1 = -2.182 \times 10^{-18}$ $= -\frac{2.182 \times 10^{-18}}{1.6 \times 10^{-19}} = 13.6$ eV	[2]Correct E_1 in *J*. **– 2.182 × 10^{-18} J** Accept ONE mistake, OR wrong energy calculation but converted into eV correctly.	[2]Correct answer. **13.6 eV**	
2009(1) (a)	Apply a high voltage across hydrogen gas	[1] Apply voltage / put current through the gas OR Put a charge through/ heat the gas to **excite** it.		

(b)	In the Bohr model the electrons can only exist in specific energy levels. When the electron drops from a higher to a lower energy level a photon is emitted with an energy that is the difference between the two energy levels. Since there are only specific energy changes possible there are only specific energies, frequencies or wavelengths possible. The Rutherford model did not have any energy levels, so it could not explain the specific wavelengths of light.	[1]Clear reference to energy levels OR Electrons moving between shells/levels create/absorb photons/spectra/light	[1]Description of change to lower **energy** level producing light/photons(not just spectra) OR [reference to energy levels AND only certain wavelengths/frequencies/ energies of photon possible] OR only certain electron energy level changes possible	[1]Only specific energy levels are possible so **only specific changes possible** AND photons created by a drop in energy levels THEREFORE **only certain wavelengths/energies of photons possible** Accept equivalent explanation of absorption spectra.
(c)	$E_n = -\frac{hcR}{n^2}$ $= \frac{6.63 \times 10^{-34} \times 3 \times 10^8 \times 1.097 \times 10^7}{2^2}$ $= -5.455 \times 10^{-19}$ J	[2] Correct answer.		

(d)	Because the line is in the visible part of the spectrum we know that n = 2 at the lowest level. Method (a): $E_{photon} = 2.86 \times 1.60 \times 10^{-19}$ $= 4.576 \times 10^{-19}\,J$ $f = \frac{E}{h} = \frac{4.576 \times 10^{-19}}{6.63 \times 10^{-34}} = 6.90196 \times 10^{14}\,Hz$ $\lambda = \frac{v}{f} = \frac{3.00 \times 10^{8}}{6.90196 \times 10^{14}} = 4.34659 \times 10^{-7}\,m$ $\frac{1}{\lambda} = R\left(\frac{1}{S^2} - \frac{1}{L^2}\right) = R\left(\frac{1}{2^2} - \frac{1}{n^2}\right)$ $\frac{1}{n^2} = \frac{1}{2^2} - \frac{1}{\lambda R} = \frac{1}{4} - \frac{1}{4.34659 \times 10^{-7} \times 1.097 \times 10^{7}}$ $= 0.040278$ $n = \frac{1}{\sqrt{0.040278}} = 4.9827$ $= 5$ Method (b): $E_2 = -5.45 \times 10^{-19}\,J$ $E_{photon} = 2.86 \times 1.60 \times 10^{-19}$ $= 4.576 \times 10^{-19}\,J$ $E_{final} = E_2 + E_{photon}$ $= -5.45 \times 10^{-19} + 4.576 \times 10^{-19}$ $= -8.7883 \times 10^{-20}\,J$ $E_n = -\frac{hcR}{n^2}$ $n = \sqrt{\frac{-hcR}{E_n}}$ $= \sqrt{\frac{-6.63 \times 10^{-34} \times 3.00 \times 10^{8} \times 1.097 \times 10^{7}}{-8.7883 \times 10^{-20}\,J}}$ $= 4.9827$ $= 5$	[2]Used ΔE instead of E_n for method (b) to get n=2.18	[2]Correct calculation of $\lambda = 434 \times 10^{-9}$ OR correct method (a) with wrong λ OR correct method (b) but makes a mistake with a signs in the energy calculation. This gives a value of **n= 1.47** OR Uses correct method but has n=1 for initial state. This gives a value of **n= 1.125**	[2]Correct calculation of n with correct working

2008(2) (a)	$v = f\lambda \Rightarrow \lambda = 3.00 \times 10^{8} \div 6.250 \times 10^{14}$ $= 4.8000 \times 10^{-7}$ m = $\mathbf{4.80 \times 10^{-7}}$ **m**	[2]Correct answer. [1]Answer rounded to 3sf plus 3 correct units given.		
(b)	UV photons have less energy than X-ray photons, but need more energy than visible photons. Energy cannot be created or destroyed / conservation of energy. So a UV photon cannot provide enough energy to form an X-ray photon, but can provide enough energy to form a visible photon.	[1]Links energy conservation / quantum nature of light concept to visible OR X-ray situation.	[1]Links energy conservation concept and quantum nature of light concept to visible AND X-ray situations.	
(c)	$\Delta E = E_f - E_i \Rightarrow E_f = E_i + \Delta E$ $= -8.24 \times 10^{-20} + -4.144 \times 10^{-19}$ J $= -4.968 \times 10^{-19} =$ $\mathbf{-4.97 \times 10^{-19}}$ **J**	[2]Correct answer consistent with incorrect handling of +/–.	[2]Correct answer.	
(d)	$E_{heat} = E_{UV} - E_{light}$ $= hf_{UV} - 4.144 \times 10^{-19}$ $= 6.63 \times 10^{-34} \times 3.86 \times 10^{15} - 4.144 \times 10^{-19}$ = $\mathbf{2.55918 \times 10^{-18}}$ $- 0.4144 \times 10^{-18}$ $= 2.14478 \times 10^{-18}$ = $\mathbf{2.14 \times 10^{-18}}$ **J**	[2]Correct value for E_{UV}.	[2]Correct answer.	

(e)	For the T-shirt to look white, the frequencies, and hence energies of the photons emitted from the phosphor, must produce the necessary colours that sum to white. This means the phosphor must have electron energy levels with energy values that have differences that give the required set of values.	[1]One key idea identified.	[1]Two key ideas identified and linked.	[1]Key ideas identified and linked are: Frequencies of emitted photons must produce the colours that add up to white. Frequency of a photon depends on its energy. Number and values of phosphor electron energy levels must allow this.
2007(2) (a)	n = 2	[1]Correct level.		
(b)	When white light is shone through hydrogen gas, the photons that have energy values that exactly coincide with one of the energy differences between the allowed energy levels for hydrogen will be absorbed by the hydrogen electron. This means the light frequency related to this energy will be removed. Because red light has lowest frequency and hence lowest energy in the visible spectrum, the transition that involves red light must involve the least energy difference and so is between levels 2 and 3. Because this frequency of light is removed, there is a dark line.	[1]Idea of energy being absorbed from photons and so removing that frequency. OR Absorption of specific frequency /wavelength (implied) associated with a transition.	[1]Clear and correct idea of energy being absorbed by electrons from specific photons and so removing that specific frequency.	[1]Clear and correct idea of energy being absorbed from specific photons and so removing that specific frequency. Correct transition clearly explained.

(c)	$\Delta E = hf = (13.6 - 1.51)$ eV $= 12.09 \times 1.6 \times 10^{-19}$ J $\Rightarrow f = \frac{1.9344 \times 10^{-18}}{6.63 \times 10^{-34}}$ $= 2.91765 \times 10^{15} = \mathbf{2.92 \times 10^{15}\ Hz}$		[2]Correct answer. (Note: 2.93 if used $E = hcR/n^2$ or Rydberg formula)	
(d)	$\Delta E(4 \rightarrow 1) = hcR\left(\frac{1}{1^2} - \frac{1}{4^2}\right)$ $= 2.0456 \times 10^{-18}$ $\Delta E(n \rightarrow 1) = hf = 6.63 \times 10^{-34} \times 3.200 \times 10^{15}$ $= 2.1216 \times 10^{-18}$ J $\Delta E(4 \rightarrow n) = (2.1216 - 2.04556) \times 10^{-18}$ $= 0.07604 \times 10^{-18}$ J $\Delta E(4 \rightarrow n) = hf \Rightarrow f = \frac{0.07604 \times 10^{-18}}{6.63 \times 10^{-34}}$ $= 1.14691 \times 10^{14} = \mathbf{1.15 \times 10^{14}\ Hz}$	[1]A correct answer implies knowledge of concepts. [2]Correct intermediate energy level calculated OR correct energy of first photon ($= 0.07604 \times 10^{-18}$)		[2]Correct answer. (Note: 1.14 if found intermediate energy level =6, incorrect if fractional energy levels used)
2006(2) (a)	$\frac{1}{\lambda} = 1.097 \times 10^7\left(\frac{1}{2^2} - \frac{1}{3^2}\right) = 1.5236 \times 10^6$ $\Rightarrow \lambda = 6.563 \times 10^{-7}$ m	[2]Correct answer. [1]Answer given to 4 sf plus 4 answers given with a (correct) unit.		

2(b)	Light of this frequency is produced when an electron in the third energy level (second excited state) falls to the second energy level. The energy lost is then released as a photon of light, with energy $E = hf$. OR consistent with transition from 2(a).	[1]Recognition that: energy transition from 3rd to 2nd level is required / loss in energy during transition creates a photon.	[1]Correct energy transition linked to photon / specific frequency production.	

(c)	Energy in an energy level is $E_n = \frac{-hcR}{n^2}$ Energy lost = $\frac{hcR}{1^2} - \frac{hcR}{7^2}$ $= 6.63 \times 10^{-34} \times 3.00 \times 10^8 \times 1.097 \times 10^7 \times (1 - \frac{1}{49}) = 2.13740 \times 10^{-18}$ J Energy 1st photon = hf $= \frac{hc}{\lambda} = 6.63 \times 10^{-34} \times \frac{3.00 \times 10^8}{2.165 \times 10^{-6}} = 9.18707 \times 10^{-20}$ J $\Rightarrow$ Energy 2nd photon = $2.13740 \times 10^{-18} - 9.18707 \times 10^{-20} = 2.0455 \times 10^{-18}$ J λ from $E = hf = \frac{hc}{\lambda}$ so $\lambda = \frac{hc}{E}$ $= 6.63 \times 10^{-34} \times \frac{3.00 \times 10^8}{2.0455 \times 10^{-18}}$ $= 9.7238 \times 10^{-8} = 9.72 \times 10^{-8}$ m OR $\frac{1}{2.165x10^{-6}} = 1.097x10^7\left(\frac{1}{s^2} - \frac{1}{7^2}\right)$ $s = 4$ $\frac{1}{\lambda} = 1.097x10^7\left(\frac{1}{1^2} - \frac{1}{4^2}\right)$ $\lambda = 9.7235x10^{-8} m$	[2]Correct wavelength consistent with incorrectly calculated energy of 2nd photon **OR** Intermediate energy level 4 **OR** Correct wavelength for incorrect intermediate energy level. **OR** Total energy change correct.	[2]Correct energy of 2nd photon.	[2]Correct answer.

(d)	Some wavelengths are missing because the photons with those wavelengths must have been absorbed. This absorption must have occurred because there are atoms between the sun and the Earth that have electrons whose energy levels have transitions that correspond to the energy of the photons that are missing.	[1]Correct idea of absorption of light by electrons/ atoms/ elements.	[1]Recognition that specific wavelengths of light are absorbed by electrons in atoms/ absorbed for electron transition.	[1]Correct link to absorption of photons of specific wavelengths due to allowed electron transitions.
2005(2) (a)	*X* will be in ultraviolet region.	Correct statement.		
(b)	$\frac{1}{\lambda} = R\left(\frac{1}{S^2} - \frac{1}{L^2}\right)$ $= 1.10 \times 10^7 \times \left(\frac{1}{1^2} - \frac{1}{2^2}\right)$ $\Rightarrow \lambda = 1.21 \times 10^{-7}$ m		Correct answer. (Accept energy difference calculation.)	
(c)	Lines in the visible part of the spectrum are from transitions to the n = 2 level. Red light has a low frequency and so the energy difference between the levels of the transition must be low. A transition from the level immediately above will involve the least energy difference and so the red line is produced from an electron transition from the n = 3 to n = 2 level.	ONE correct and relevant statement: from n = 3 to n = 2 / energy difference between the levels must be low / red line is produced from the least energy transition.	Visible is jump to n=2 and Link made between the low frequency of red light and the need for a low energy difference between the levels.	Explanation is clear, concise and accurate – clear understanding of link between smallest energy gap and smallest frequency (longest wavelength) photon.

(d)	Energy in ground state $= \frac{-hcR}{n^2}$ $= -2.1879 \times 10^{-18} \quad = -2.19 \times 10^{-18}$ J $(= -13.7$ eV$)$ (negative not required.)	Correct answer. Rounded to 3 sig fig plus three answers given with correct unit.		
(e)	$E_3 = \frac{-hcR}{3^2}, \quad E_5 = \frac{-hcR}{5^2}$ $E_3 - E_5 = -hcR(\frac{1}{3^2} - \frac{1}{5^2})$ $= -1.56 \times 10^{-19}$ J Photon energy $= 1.56 \times 10^{-19}$ J		Correct answer, ignore sign.	