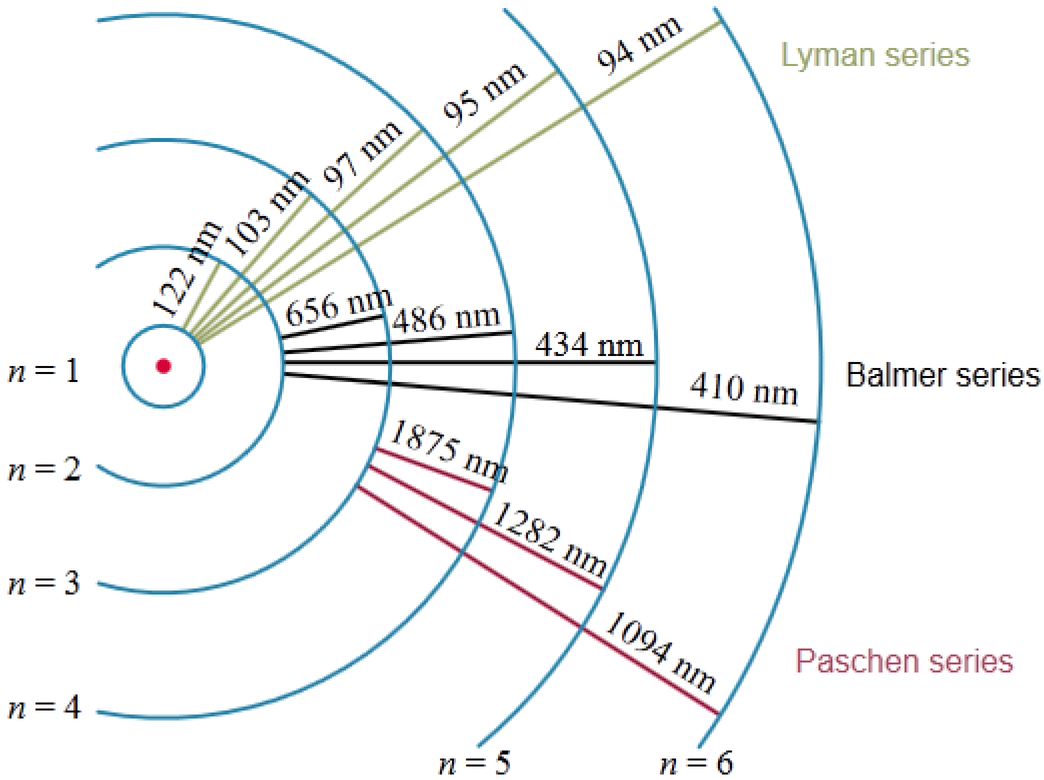
ATOMS: SPECTRA QUESTIONS

SODIUM LAMPS (2012;2)

(d) In 1802 William Wollaston noted the appearance of dark lines in the spectrum of sunlight. These lines are due to the presence of certain chemical elements in gases surrounding the Sun.
- (i) Explain why sharp dark lines appear only at specific wavelengths.
- (ii) Explain how a comparison between the spectrum of sunlight, with the dark lines, and the spectrum of light from a sodium lamp, can identify that sodium is one of the elements in the Sun's atmosphere.

ATOMIC SPECTRA (2010;1)

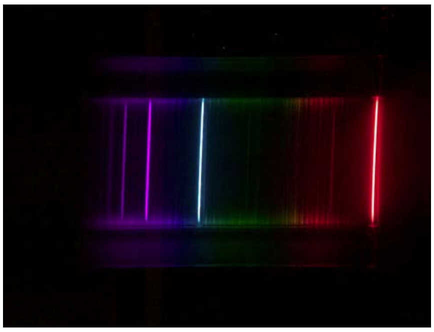


(a) The above diagram illustrates the Bohr model of the hydrogen atom.
- (a) Explain what the circles and lines represent.
- (b) Explain how this model explains emission line spectra **and** absorption line spectra caused by hydrogen.

(b) The line between n = 5 and n = 2 is labelled 434 nm (4.34×10^{-7} m). Show that this is correct for a hydrogen atom.

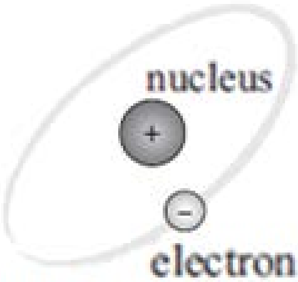
(c) Calculate the energy of a hydrogen atom in its ground state. Express your answer in eV.

THE HYDROGEN SPECTRUM (2009;1)



(a) Briefly describe how the light that formed the hydrogen line spectrum illustrated could be produced experimentally.

(b) Rutherford's model pictured an atom with 'electrons orbiting a solar nucleus', as illustrated below. The Bohr model was proposed to explain the line spectra of hydrogen, something that the Rutherford model could not explain.

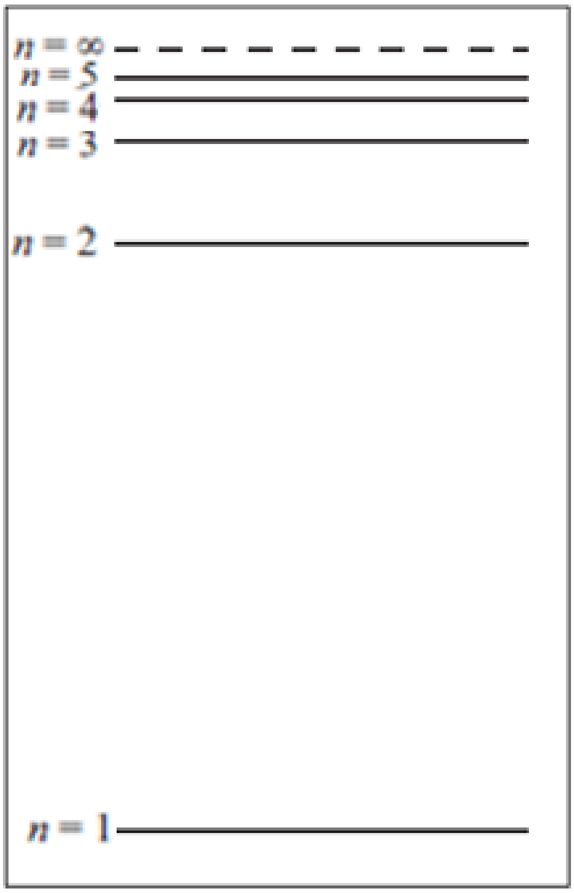


The Rutherford model for a hydrogen atom.

Discuss how the Bohr model was able to explain the line spectrum that the Rutherford model could not.

(c) The hydrogen spectrum in the visible region is part of the Balmer series. Energy transitions that give rise to this series are to the n = 2 level (emission) and from the n = 2 level (absorption). Calculate the energy of the n = 2 level.

(d) The diagram shows energy levels in a hydrogen atom. A continuous spectrum from a star shows absorption lines in the visible part of the spectrum. One dark line indicates the absorption of photons with an energy of 2.86 eV. An astronomer believes that this is due to the presence of hydrogen atoms surrounding the star. Calculate the final energy level, n, when a photon of light of this frequency is absorbed by an atom of hydrogen.

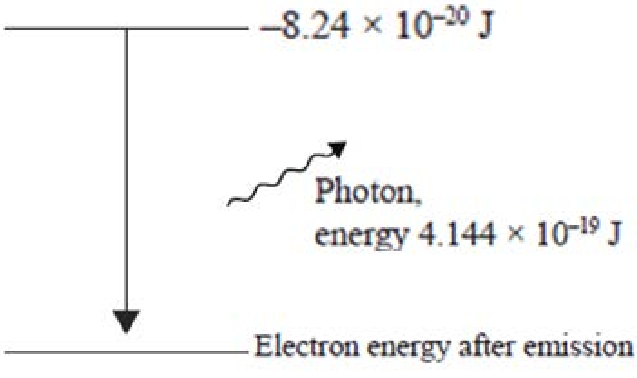


ENERGY LEVELS (2008;2)

Planck's constant $= 6.63 \times 10^{-34}$ J s
Speed of light $= 3.00 \times 10^{8}$ m s^{-1}

A 'black' lamp produces photons that are in the ultraviolet (UV) part of the electromagnetic spectrum. When UV is absorbed by some materials, it makes them glow in the dark. These materials are called phosphors. The white T-shirt in the photo glows because the detergent it was washed in contains phosphors that remain in the fabric after laundering. When an electron in a phosphor is excited by a UV photon, it comes down from its excited state by emitting a visible photon. Some of the original UV energy is retained in the phosphor as thermal energy.
An electron is excited by a UV photon, causing a visible photon of frequency 6.250×10^{14} Hz and energy 4.144×10^{-19} J to be emitted.

(a) Calculate the wavelength of the light photon. Round your answer to the correct number of significant figures.
(b) Suggest a reason why the photons emitted from the phosphor are in the visible region of the electromagnetic spectrum but not in the X-ray region.



(c) The visible photon is emitted when the excited electron drops to a lower energy level. If the excited electron has energy of -8.24×10^{-20} J, calculate the energy of the electron after the photon has been emitted.
(d) The UV photon whose energy was used to excite the electron has a frequency of 3.86×10^{15} Hz. Calculate the amount of heat energy gained by the phosphor after one absorption-emission event.
(e) The colour of the T-shirt under the 'black lamp' light is white. White light is seen when light of many different wavelengths is present. Explain what this tells you about the energy levels in the phosphor material.

HYDROGEN SPECTRUM (2007;2)

Speed of light	$= 3.00 \times 10^{8}$ m s^{-1}
Charge on the electron	$= 1.6 \times 10^{-19}$ C
Planck's Constant	$= 6.63 \times 10^{-34}$ J s
Rydberg Constant	$= 1.097 \times 10^{7}$ m^{-1}

The electron in the hydrogen atom emits or absorbs electromagnetic radiation when it moves between different energy levels. The visible part of the spectrum emitted by hydrogen can be seen in the laboratory by applying a high voltage to a hydrogen discharge tube. The diagram below represents some of the electron energy levels in the hydrogen atom.

(a) To which energy level does the electron drop when it emits visible light?

(b) The absorption spectrum for hydrogen gas consists of a series of dark lines within the full spectrum of colours. Explain clearly how the dark line in the red part of the spectrum is produced.

(c) Calculate the frequency of the photon produced when an electron drops from the second excited state (n = 3) to the ground state (n = 1).

An electron in energy level 4 jumps to a higher energy level, and then drops down to the ground state, releasing a photon of frequency 3.200×10^{15} Hz.

(d) Calculate the frequency of the photon required for the first jump.

SOLAR POWER (2006;2)

Rydberg's constant $= 1.097 \times 10^{7}\ \text{m}^{-1}$
Planck's constant $= 6.63 \times 10^{-34}\ \text{J s}$
Speed of light $= 3.00 \times 10^{8}\ \text{m s}^{-1}$

Nuclear reactions in the Sun produce light. The main element in the Sun is hydrogen. The spectrum of hydrogen can be observed in the laboratory with a hydrogen discharge tube. The visible lines in the hydrogen spectrum are called the Balmer series and are described by the formula:

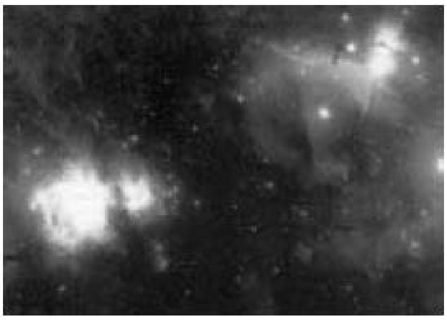
$$\frac{1}{\lambda} = R\left(\frac{1}{S^2} - \frac{1}{L^2}\right)$$

where S =2.

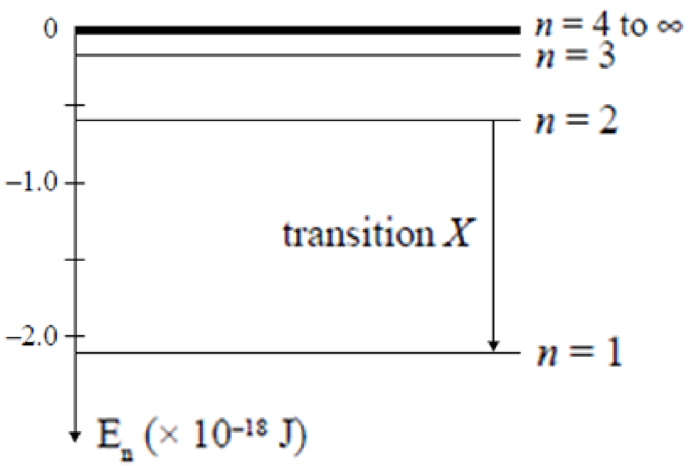
(a) Calculate the wavelength of the lowest frequency line in the Balmer series (L = 3). Give the answer to the correct number of significant figures.
(b) Explain how light of this particular frequency is produced in the hydrogen atom.
(c) An electron in the 6th excited state (L = 7) returns to the ground state in two jumps. It releases one photon with a wavelength of 2.165×10^{-6} m. What is the wavelength of the second photon?
(d) The Sun emits all wavelengths. However, when a solar spectrum is observed on Earth, it contains black lines that correspond to missing wavelengths. Give an explanation, in terms of energy absorption by electrons, for why some of the wavelengths of light in the solar spectrum are missing when the light reaches Earth.

LIGHT EMITTED FROM A STAR (2005;2)

Light from stars is photons of electromagnetic radiation created by electron transitions between energy states. These photons produce spectra that identify the atoms that are producing the light. A common element that produces light from a star is hydrogen.


The Horsehead Nebula has a pink glow that comes from ionised hydrogen gas.

The possible energy states (levels) of the hydrogen atom electron are shown in the diagram below.



(a) In which part of the electromagnetic spectrum is the radiation emitted by transition X?

(b) Calculate the wavelength of the photons emitted by transition X.
(c) Explain which transition produces the red line in the visible part of the hydrogen atom
(d) Calculate the value of the ground state energy of the hydrogen electron. Give your answer to the correct number of significant figures.
(e) In order for an electron in a hydrogen atom to move from the third energy level to the fifth energy level, a photon of electromagnetic radiation must be absorbed. Calculate the energy of this photon.

QUESTION TWO (2004;1)

Use the following information when answering this question:

$$\text{Speed of light} = 3.00 \times 10^{8}\ \text{m s}^{-1}$$
$$\text{Planck's constant} = 6.63 \times 10^{-34}\ \text{J s}$$
$$\text{Rydberg's constant} = 1.10 \times 10^{7}\ \text{m}^{-1}$$

Light energy is produced when the excited electrons of atoms drop from higher energy levels to lower energy levels.

(a) The electron in a hydrogen atom falls from the 5th energy level to the 2nd energy level.
  (i) Calculate the energy of the electron at the 2nd energy level. Give your answer to an appropriate number of significant figures.
  (ii) Calculate the wavelength of the light emitted by this electron transition.

The hydrogen atom, whose electron is in the 2nd energy level, is now ionised.

(b) Calculate the minimum frequency of the photon that can ionise the hydrogen atom.

In a hydrogen atom, electron transitions to the 2nd energy level produce visible light.

(c) Which transition will produce visible light of the lowest frequency? Explain your answer.

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)	$\frac{1}{\lambda} = \mathrm{R}\left(\frac{1}{S^2} - \frac{1}{L^2}\right)$ $\frac{1}{\lambda} = 1.097 \times 10^7 \left(\frac{1}{2^2} - \frac{1}{5^2}\right)$ $= 2.304 \times 10^6$ $\lambda = 4.341 \times 10^{-7}$ m	[2]Correct method.		
(c)	$E_\mathrm{n} = \frac{\mathrm{hcR}}{n^2}$ $E_1 = -\frac{6.626 \times 10^{-34} \times 3.00 \times 10^8 \times 1.097 \times 10^7}{1^2}$ $E_1 = -2.182 \times 10^{-18}$ $= -\frac{2.182 \times 10^{-18}}{1.6 \times 10^{-19}} = 13.6$ 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): $E_{photon} = 2.86\times1.60\times10^{-19}$ $= 4.576\times10^{-19}\,J$ $f = \frac{E}{h} = \frac{4.576\times10^{-19}}{6.63\times10^{-34}} = 6.90196\times10^{14}\,Hz$ $\lambda = \frac{v}{f} = \frac{3.00\times10^{8}}{6.90196\times10^{14}} = 4.34659\times10^{-7}\,m$ $\frac{1}{\lambda} = R\left(\frac{1}{S^2} - \frac{1}{L^2}\right) = R\left(\frac{1}{2^2} - \frac{1}{n^2}\right)$ $\frac{1}{n^2} = \frac{1}{2^2} - \frac{1}{\lambda R} = \frac{1}{4} - \frac{1}{4.34659\times10^{-7}\times1.097\times10^{7}}$ $= 0.040278$ $n = \frac{1}{\sqrt{0.040278}} = 4.9827$ $= 5$ Method (b): $E_2 = -5.45\times10^{-19}\,J$ $E_{photon} = 2.86\times1.60\times10^{-19}$ $= 4.576\times10^{-19}\,J$ $E_{final} = E_2 + E_{photon}$ $= -5.45\times10^{-19} + 4.576\times10^{-19}$ $= -8.7883\times10^{-20}\,J$ $E_n = -\frac{hcR}{n^2}$ $n = \sqrt{\frac{-hcR}{E_n}}$ $= \sqrt{\frac{-6.63\times10^{-34}\times3.00\times10^{8}\times1.097\times10^{7}}{-8.7883\times10^{-20}\,J}}$ $= 4.9827$ $= 5$	[2]Used ΔE instead of E_n for method (b) to get n=2.18	[2]Correct calculation of $\lambda = 434\times10^{-9}$ 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.	