inkPhysics

QUESTION WORKBOOK

Energy levels and photon emission

Quantum phenomena · Year 12

AQA 3.2.2.3 · CIE 22.4.1, 22.4.2, 22.4.3
OCR A 5.5.2(a), 5.5.2(b), 5.5.2(c), 5.5.2(d), 5.5.2(e), 5.5.2(f)

19 questions · 57 marks

NAVY SCREEN EDITION

SECTION A · Warm-up

A1 Explain what the sharp lines of an emission line spectrum tell us about the energy of electrons in an atom. [2]

A2 An electron transition releases 3.0 eV of energy. Calculate this in joules ($e = 1.60 \times 10^{-19}$ C). [2]

A3 State the equation relating the frequency of an emitted photon to the energy levels involved. [2]

A4 State what is meant by the ground state of an atom. [1]

A5 The energy levels of an atom are given negative values. Explain why. [2]

A6 White light is passed through a cool gas and then through a diffraction grating. Dark lines appear in the otherwise continuous spectrum. Explain why. [2]

SECTION B · Standard

B1 An electron falls from a level at -1.5 eV to a level at -3.4 eV. Taking $h = 6.63 \times 10^{-34}$ J s, $c = 3.0 \times 10^{8}$ m s^{-1} and $e = 1.60 \times 10^{-19}$ C, calculate (a) the photon frequency and (b) its wavelength. [4]

B2 A transition releases a photon of energy 2.55 eV. Calculate its wavelength. [3]

B3 An electron drops from a level at -0.54 eV to a level at -3.40 eV. Calculate the wavelength of the emitted photon. [3]

B4 A photon of wavelength 486 nm is emitted from an atom. Calculate the energy of the transition, in electronvolts. [3]

B5 The red line in the spectrum of atomic hydrogen has a wavelength of 656 nm and is produced by electrons falling to the level at -3.40 eV. Determine the energy of the level from which the electrons fall ($h = 6.63 \times 10^{-34}$ J s, $c = 3.0 \times 10^{8}$ m s^{-1}, $e = 1.60 \times 10^{-19}$ C). [3]

B6 A sodium atom has its ground state at -5.14 eV and its first excited level at -3.04 eV. Calculate the frequency of the photon that must be absorbed to raise an electron from the ground state to the first excited level ($h = 6.63 \times 10^{-34}$ J s, $e = 1.60 \times 10^{-19}$ C). [3]

B7 Four of the energy levels of atomic hydrogen are -0.85 eV, -1.51 eV, -3.40 eV and -13.6 eV (the ground state). A line in the ultraviolet spectrum of hydrogen has a wavelength of 122 nm. Deduce which transition produces this line ($h = 6.63 \times 10^{-34}$ J s, $c = 3.0 \times 10^{8}$ m s^{-1}, $e = 1.60 \times 10^{-19}$ C). [4]

SECTION C · Stretch

C1 An atom has levels at 0 eV (ionised), -1.5 eV and -13.6 eV. An electron falls from the -1.5 eV level to the -13.6 eV level. Calculate the wavelength of the photon emitted and state which part of the spectrum it lies in. [4]

C2 An atom has three energy levels: -1.5 eV, -3.4 eV and -13.6 eV. State how many different spectral lines can be produced by transitions between them, and calculate the wavelength of the line with the longest wavelength. [4]

C3 Explain why the line spectrum of each element is unique. [3]

C4 An atom has three energy levels: -0.9 eV, -2.8 eV and -4.6 eV (the ground state). A student claims that all three possible spectral lines lie in the visible range, 400 nm to 700 nm. Deduce whether the claim is correct ($h = 6.63 \times 10^{-34}$ J s, $c = 3.0 \times 10^{8}$ m s^{-1}, $e = 1.60 \times 10^{-19}$ C). [5]

C5 An atom has energy levels at -1.0 eV, -2.5 eV and -6.0 eV (the ground state). An electron in the -1.0 eV level returns to the ground state via the -2.5 eV level. State the number of photons emitted and calculate the wavelength of each ($h = 6.63 \times 10^{-34}$ J s, $c = 3.0 \times 10^{8}$ m s^{-1}, $e = 1.60 \times 10^{-19}$ C). [4]

C6 A hydrogen atom is in its first excited state, at -3.40 eV. Ionisation corresponds to 0 eV. Determine the maximum wavelength of electromagnetic radiation that can ionise the atom from this state ($h = 6.63 \times 10^{-34}$ J s, $c = 3.0 \times 10^{8}$ m s^{-1}, $e = 1.60 \times 10^{-19}$ C). [3]