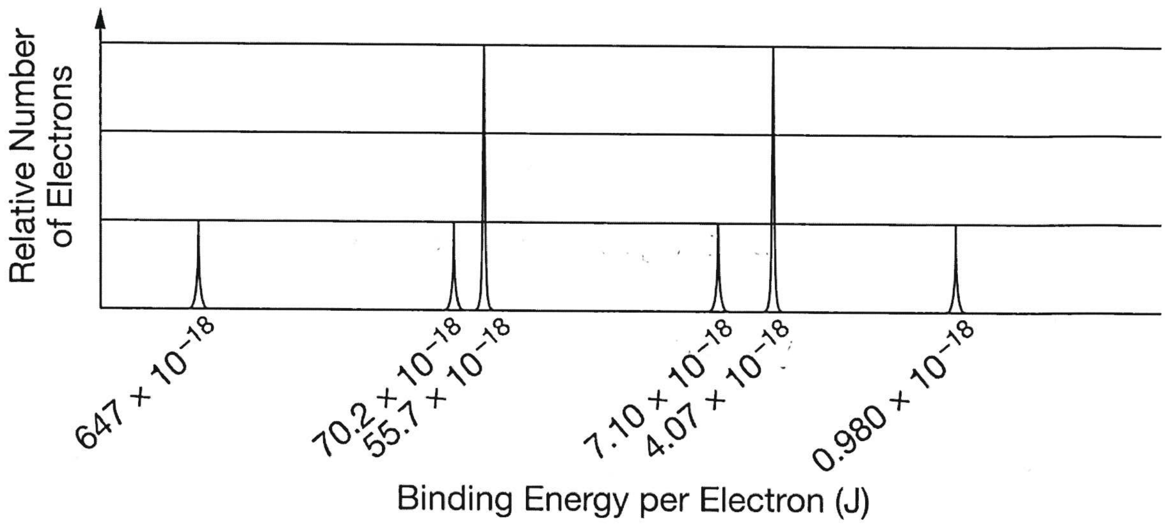
Question #1

The complete photoelectron spectrum of an element in its ground state is represented below.



(a) Based on the spectrum,

(i) write the ground-state electron configuration of the element, and $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

(ii) identify the element. Calcium

(b) Calculate the wavelength, in meters, of electromagnetic radiation needed to remove an electron from the valence shell of an atom of the element.

Valence shell = 0.980×10^{-18} J

$E = h\nu$

$0.980 \times 10^{-18} J = 6.626 \times 10^{-34} J \cdot s \ \nu$

$\nu = 1.48 \times 10^{15} \ sec^{-1}$ = frequency

$c = \lambda \nu$

$2.998 \times 10^{8} \frac{m}{sec} = \lambda \ (1.48 \times 10^{15} \ sec^{-1})$

$\lambda = 2.03 \times 10^{-7} m$ = wavelength

Question #2

A student uses visible spectrophotometry to determine the concentration of $CoCl_2(aq)$ in a sample solution. First the student prepares a set of $CoCl_2(aq)$ solutions of known concentration. Then the student uses a spectrophotometer to determine the absorbance of each of the standard solutions at a wavelength of 510 nm and constructs a standard curve. Finally, the student determines the absorbance of the sample of unknown concentration.

A wavelength of 510 nm corresponds to an approximate frequency of $6 \times 10^{14}\ s^{-1}$. What is the approximate energy of one photon of this light?

A) 9×10^{47} J

B) 3×10^{17} J

C) 5×10^{-7} J

D) 4×10^{-19} J

ν = frequency = 6×10^{14} sec^{-1}

$E = h\nu$

$E = (6.626 \times 10^{-34}\ J\,sec)(6 \times 10^{14}\ sec^{-1})$

$\longrightarrow$ $E = 3.9756 \times 10^{-19}\ J = 4 \times 10^{-19}\ J$ = Energy

NOTE: If you want Joules per mole of photons

$4 \times 10^{-19}\ J \times \frac{6.022 \times 10^{23}\ \text{photon}}{1\ \text{mole photons}} = 240880\ \frac{J}{\text{mol photons}}$

Question #3

The diagram above represents the photoelectric effect for a metal. When the metal surface is exposed to light with increasing frequency and energy of photons, electrons first begin to be ejected from the metal when the energy of the photons is 3.3×10^{-19} J. $\leftarrow$ or 2.0×10^5 J per mole of photons

Color	Wavelength
Red	647 – 760 nm
Orange	585 – 647 nm
Yellow	575 – 585 nm
Green	491 – 575 nm
Blue	424 – 491 nm
Violet	300 – 424 nm

Using the wavelength information provided above, what is the color of the light?

A) Red

B) Orange

C) Yellow

D) Blue

$\frac{2.0 \times 10^5\ J}{1\ \text{mol of photons}} \times \frac{1\ \text{mol photon}}{6.022 \times 10^{23}} = 3.3 \times 10^{-19}\ J$

$E = h\nu$

$3.3 \times 10^{-19}\ J = (6.626 \times 10^{-34}\ J \cdot sec)\ \nu$

ν = frequency = 5.0×10^{14} sec^{-1}

$c = \lambda\nu$

$2.998 \times 10^8\ \frac{m}{sec} = \lambda(5.0 \times 10^{14}\ sec^{-1})$

$\lambda = 6.0 \times 10^{-7}$ m = wavelength

$\rightarrow$ $\boxed{600\ nm}$