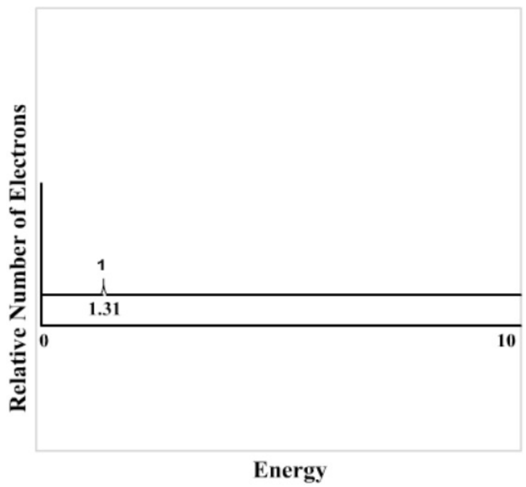
Photoelectron spectroscopy (PES) is a technique that is used to gather information about the electrons in an atom. In PES, an atom is bombarded with photons. Some of the photons are absorbed and electrons are emitted. The electrons are collected and their energy is analyzed. During PES, we know the energy of the photons with which we excite the electrons. Since energy is conserved, the difference in the energy between the photons sent into the atom and the energy of the emitted electrons will be the potential energy of the electrons when they are attached to the atom. This is the ionization energy (IE). The potential energy of the electron in the atom is the work needed to remove the electron from the atom.

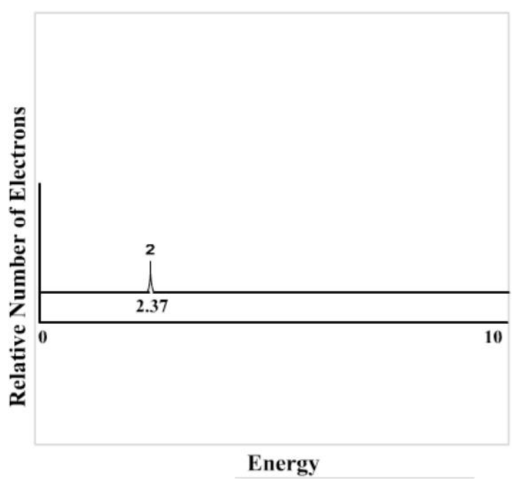
$$IE = E_{photon} - E_{emitted\ e-}$$

The following refers to the PES data for hydrogen:



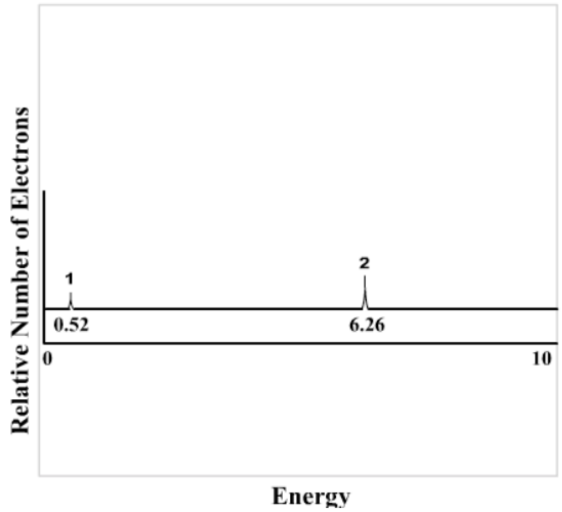
The peak on the graph has two numbers. The whole number on the top of the peak is the number of electrons the element has. The decimal number is the ionization E needed to remove the electron from the atom in MJ/mol.

Now for the PES for helium:



You'll see that the number of electrons is 2 and the IE is higher than hydrogen. That's because He has 2 electrons and H has 1. Also, H and He have the same level of shielding (none) while He has 2 protons and H has 1. So the higher nuclear charge has a stronger attraction for the electrons, meaning that it takes more energy to remove the electrons from He than from H.

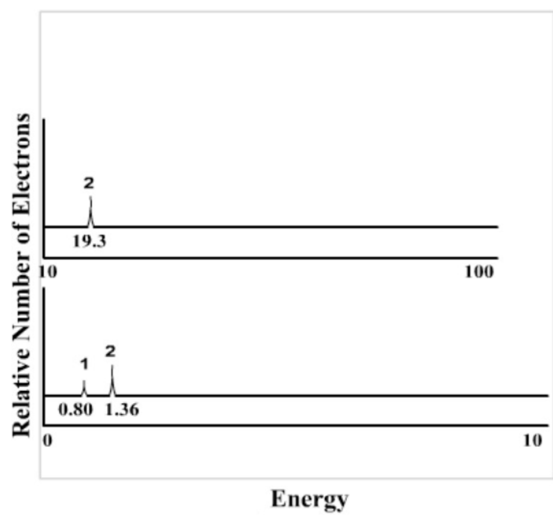
For lithium, there two sets of electrons in different energy levels.



In Li's spectrum above, there are two peaks. Of the two peaks, one is for the electron at the higher energy 2s level, and the other is for the two electrons in the lower energy 1s level. Let's think about terminology. We have said that electrons in 1s have lower potential energy than 2s. When we look at the PES spectrum, we see 1 electron at a low number, and 2 electrons at a high number. How can this be if the 2 electrons in 1s have lower energy than the 1 electron is 2s?

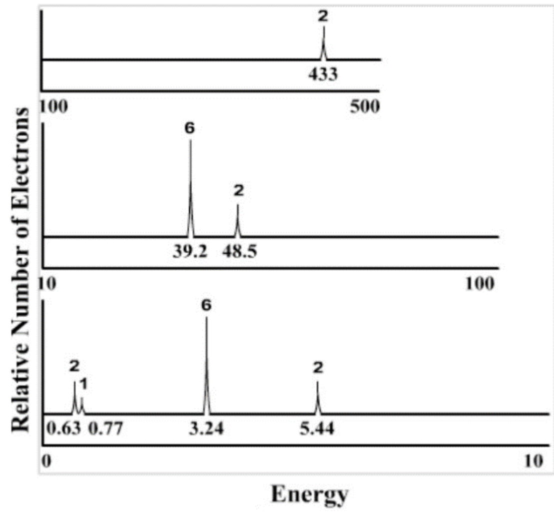
On the PES spectra, a high-energy number means the electron is closer to the nucleus and experiences more nuclear charge attraction, so more energy is needed to remove the electron. A low number means it is farther from the nucleus, experiences less nuclear charge attraction, so less energy is needed to remove it. Note that the peak for an electron farther from the nucleus is closer to the beginning of the graph while the peak for an electron closer to the nucleus is farther from the beginning of the graph.

Now look at boron's spectrum:



The two electrons at 19.3 MJ are the 1s electrons. There are also two electrons at 1.36 MJ and one electron at 0.80 MJ. The two electrons are 2s and the one electron is 2p. Why are they different? They are both in the same energy level (shell), so why do they have different potential energies? The reason the 2s and 2p orbitals have different energies is that s penetrates better than p. An s orbital can overcome the effect of shielding better than p, so more E is required to remove that 1s electron.

Looking at the spectrum for scandium:



The filling order has 3d after 4s because 3d does not penetrate as well as 4s. However, in the spectrum for scandium we can see that there are two electrons at 0.63 MJ and one at 0.77 MJ. This suggests that it is easier to remove the 4s than the 3d. This is because the third energy level electrons shield electrons in the fourth energy level, so less energy is required to remove the 4s electrons.

Questions:

1. In a photoelectron spectrum, photons of 165.7 MJ/mol strike atoms of an unknown element. If the kinetic energy of the ejected electrons is 25.4 MJ/mol, what is the ionization energy of the element?

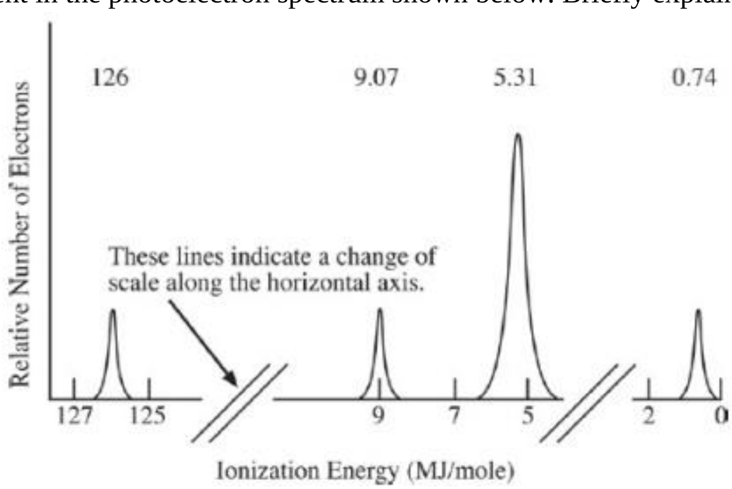
2. What determines the position and the height (intensity) of each peak in a photoelectron spectrum?

3. Why is the distance of the energy level from the nucleus important in determining the corresponding peak position in the photoelectron spectrum?

4. Based on the information provided below, draw a photoelectron spectrum for argon. Indicate the relative intensities and positions of all peaks.

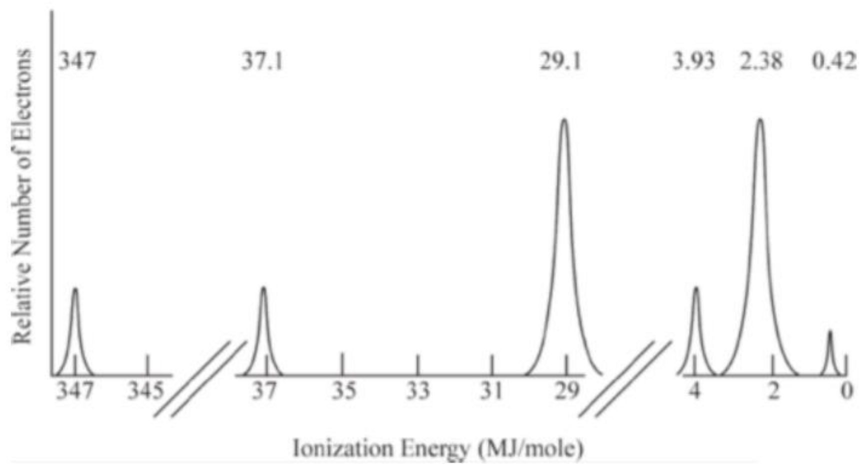
$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$
- 309.0	- 31.5	- 24.1	- 2.83	- 1.52

5. Identify the element in the photoelectron spectrum shown below. Briefly explain your reasoning.



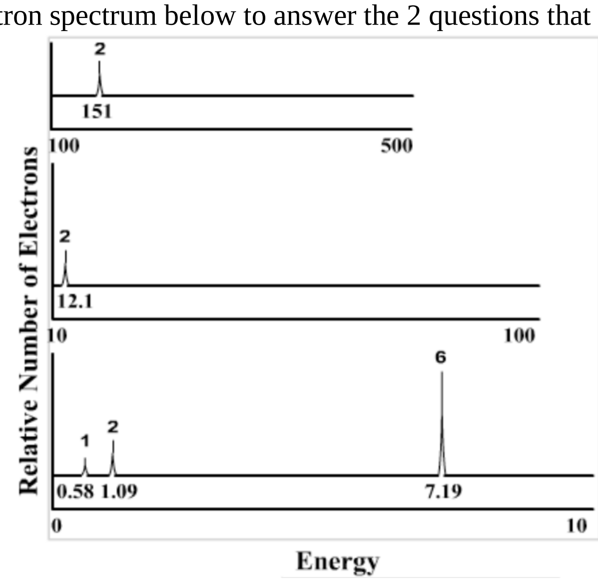
6. Identify if either of the following statements is correct. Briefly explain your reasoning:
   a) The photoelectron spectrum of Mg^{2+} is expected to be identical to the photoelectron spectrum of Ne.
   b) The photoelectron spectrum of ^{35}Cl is identical to the photoelectron spectrum of ^{37}Cl.

7. Is it possible to deduce the electron configuration for an atom from its photoelectron spectrum? If so, explain how. If not, explain why not.



8. The peak at 0.42 MJ/mol in the photoelectron spectrum above is identified as being in the 4th energy level. What element does the spectrum represent, and why?

9. Refer to the photoelectron spectrum below to answer the 2 questions that follow.



   a) The spectrum above is for which element? Explain your reasoning.
   b) Write the electron configuration for this element.

10. The threshold frequency for sodium metal (i.e., the minimum frequency that will cause the photoelectric effect) is 5.56×10^{14} s^{-1}. What is the minimum energy of a photon that will cause the effect in sodium, and will red light cause the photoelectric effect in sodium metal? Why or why not?