

# General information about the electromagnetic spectrum, electron configurations, and other related topics

## 6.1 The Wave Nature of Light

- a) Electromagnetic energy with a short wavelength has a (high/low) frequency.  
Answer\_\_\_\_\_
- b) Light with a short wavelength has a (high/low) energy.  
Answer\_\_\_\_\_
- c) Which color of visible light has the lowest energy?  
Answer\_\_\_\_\_
- d) Which light has slightly less energy than green light?  
Answer\_\_\_\_\_
- e) Radio waves can be heard by the human ear(true/false)?  
Answer\_\_\_\_\_
- f) What type of radiation has a wavelength just slightly longer than red light?  
Answer\_\_\_\_\_
- g) The term for the distance between 2 consecutive peaks in a wave is:  
Answer\_\_\_\_\_
- h) The portion of electromagnetic radiation with the wavelength approximately 400 to 700 nm is: Answer\_\_\_\_\_

1. Calculate the frequency of light that has a wavelength of 75.0 nm. What type of radiation is this?
  
  
  
  
  
  
  
  
  
  
2. A photon of visible light has a energy of  $3.4 \times 10^{-19} \text{J}$ . What is the wavelength of this light? What is the frequency of the light? What region of the EM spectrum would one find this light?

3. Calculate the energy of a photon that has a frequency of 2.5 THz.

## 6.2 Quantized Energy and Photons

### Hot objects and the quantization of energy:

#### Terms and constants:

- Planck constant [h] is  $6.626 \times 10^{-34} \text{ J}\cdot\text{s}$
- $\Delta E = h\nu$
- A quantum is a fixed amount of energy that can be emitted or absorbed as EM radiation. This varies with the frequency.

1. A photon has a frequency ( $\nu$ ) of  $2.68 \times 10^6$  Hz. Calculate its energy.
2. Calculate the energy (E) and wavelength ( $\lambda$ ) of a photon of light with a frequency ( $\nu$ ) of  $6.165 \times 10^{14}$  Hz.
3. Read this experiment and analysis for a great example of practical Planck. Just reading, no doing. <http://www.scienceinschool.org/2014/issue28/planck>

### Photo electric effect

#### Terms and constants:

- A photon is a particle EM energy that is quantized. We often use the symbol  $h\nu$  for the photon.
- The wavelength and frequency relate the energy of one particle.
- The energy of the photon = the energy absorbed by the electron = E to do work + E to travel
- $\Delta E_{\text{work}}$  is the work function ( $\phi$ ) and is also called the threshold energy, the energy needed to overcome the potential energy binding the electron to the surface of the metal.
  - The frequency related to the work function is often referred to as the *threshold frequency*.
- Any additional energy is given over to kinetic energy of the ejected electron.
- The equation can be expressed in a number of ways; Here are two most common:
  - $\Delta E_{\text{hv}} = \Delta E_{\text{work}} + \Delta E_{\text{KE}}$
  - $\Delta E_{\text{hv}} = \phi + K$

1. For biological organisms, more damage is done to cells by standing in front of a very weak, low power beam of x-rays than in front of a much brighter red light. Use the photoelectric effect to explain your answer. (Look at the handout on the website, CH 6 electrons in the atom)
2. What is the threshold frequency that can cause the release of electrons from a metal with a work function of 2.8 eV?  $1 \text{ eV} = 1.6099 \times 10^{-19} \text{ J}$ .
3. Gold has a work function of 4.82 eV. A block of gold is illuminated with ultraviolet light with  $\lambda = 160.0 \text{ nm}$ . Find the maximum kinetic energy of the emitted photoelectrons in electron volts. Find the threshold frequency for gold.

### 6.3 Line Spectra and the Bohr Model

- Bohr proposed that an electron moves in fixed orbits. The electron moves from one orbit to another when the atom absorbs or emits a photon whose energy equals the difference in energy levels (orbits).
- Bohr's model predicts only the spectrum of the H atom and other one-electron species. Bohr predicted that the atom's energy is quantized.

1. Lowest possible energy state of an atom is:  
Answer \_\_\_\_\_
2. According to Bohr, an electron in the Hydrogen atom moves around the nucleus in circular paths called: Answer \_\_\_\_\_
3. What term means only certain values of energy are allowed?  
Answer \_\_\_\_\_
4. Which type of light produces a continuous spectrum?  
Answer \_\_\_\_\_
5. Do any two atoms produce the same line spectrum?  
Answer \_\_\_\_\_
6. Calculate the wavelength of EM radiation emitted when an electron returns from an excited state of  $n = 3$  to the ground state of  $n = 1$ .
7. Calculate the wavelength of EM radiation absorbed when an electron is excited from  $n = 3$  to  $n = 7$ .

## 6.4 The Wave Behavior of Matter

**De Broglie wavelength:** Louis De Broglie hypothesis shows that the dual nature of light is a fundamental principle of all matter. When light travels through space it has wave properties; when it exchanges energy with matter, it has particle properties. De Broglie proposed that 'if a photon has wave and particle properties (frequency and wavelength) then a material particle should also have wave properties. De Broglie's theories laid the foundation for the development of the electron microscope and electron diffraction.

1. How was the wave behavior of matter used to determine the structure of DNA? (see 6.49 and 6.50)

**Heisenberg's Principle:** the Uncertainty Principle proposed by Werner Heisenberg describes the inability to determine both the position and the velocity of a particle. This is inherent to the dual nature of matter. If we have a series of objects ranging from visible and large to very small, we use different techniques to study the various objects. Large objects can be measured with a ruler; smaller objects require more magnification and resolution and use light to study the object. Small particles are affected by the introduction of light. The particle is disturbed, and moves away. We know either the position, or the velocity, but not both simultaneously. The Uncertainty Principle gives rise to the idea of probability of finding a particle in a region of space, instead of the exact location.

## 6.5 Quantum Mechanics and Atomic Orbitals

Quantum Mechanics and the math to express the energies of the electron in hydrogen atom are very complex, and past the scope of our class. We can think of the electron's energy like a standing wave. This is the approach that Schrödinger used. His equation looks simple,  $\mathcal{H}\psi = E\psi$ , Where H (for Hamiltonian operator) stands for a series of 'equations' that when solved for a given  $\psi$ , yields an allowed energy state of an atom, and  $\psi$  stands for the **wave function**. Each solution of the equation gives a state that has an associated atomic orbital. This represents the total energy of the particle (both PE and KE). The more energy we give to the standing wave, the more it will oscillate back and forth. A specific  $\psi$  is called an **orbital**. The **wave function** gives no information about the path that the electron is travelling. The position of the electron at any given time is unknown. Heisenberg's principle concludes that the more we know about the position of the electron, the less we know about the momentum of the electron. Based on the figure below, the particle spends most of its time in the lowest energy. The *problem* with the wave function? It is hard to visualize the physical meaning; instead  $\psi^2$  is used to indicate the probability of finding an electron near a point in space.  $\psi^2$  is called the **probability function**, and gives rise to **electron density diagrams or electron cloud depictions**. Solving Schrodinger equation gives rise to quantum numbers that describe specific properties of the orbitals.

- The atomic orbital  $\psi$ , is a mathematical description of the electrons wavelike behavior in an atom. The Schrödinger equation converts each allowed wave function to one of the atom's energy states
- The probability density of find the electron at a location is given by  $\psi^2$ .
- An atomic orbital is described by three quantum numbers size ( $n$ ), shape ( $l$ ), and orientation ( $m$ );  $n$  limits  $l$  to  $n-1$  values and  $l$  limits  $m$  to  $2l+1$  values
- A sublevel with  $l=0$  has a spherical orbital (s); a sublevel with  $l=1$  has two-lobed orbitals with three orientations, and a sublevel with  $l=2$  has a multi-lobed with 5 orientations. In the special case of hydrogen, the energy levels depend only on the  $n$  value.

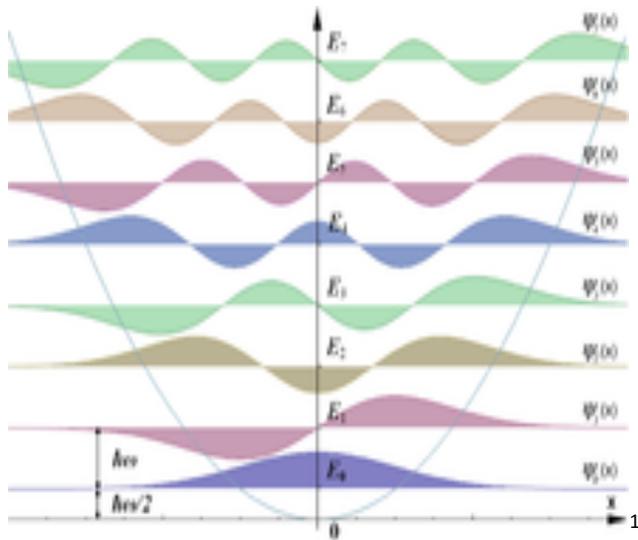


Figure 1: Wave functions for the first 8 energy levels of hydrogen.

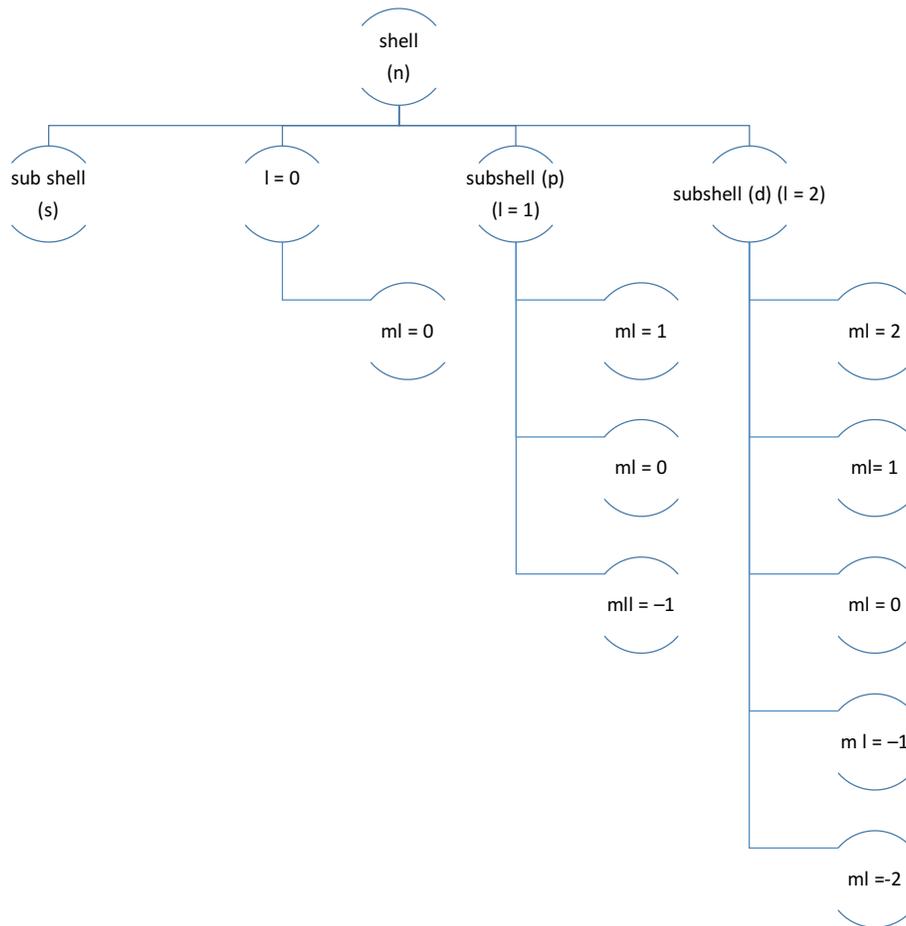
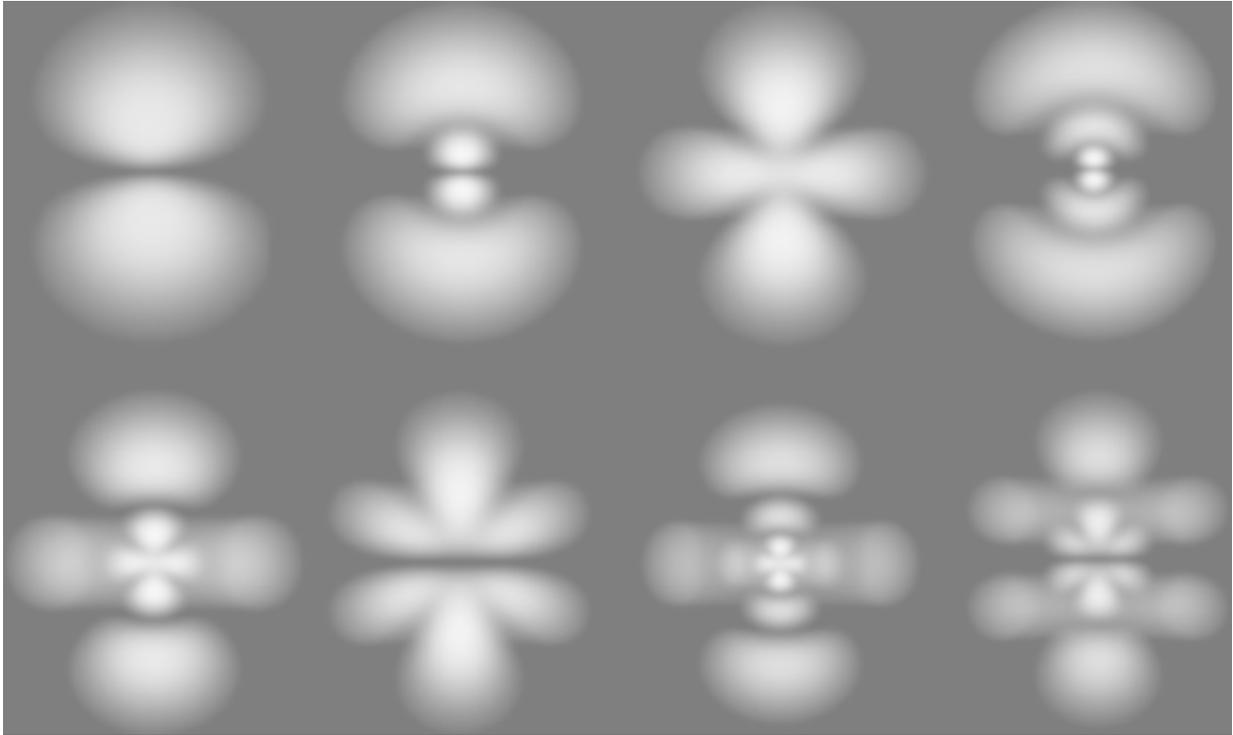
Looking at Figure 1, what is the relationship between the number of nodes and the energy of the standing wave?

### Quantum numbers:

1. The **principal quantum number**:
  - a. Use the symbol ( $n$ )
  - b. Indicates relative size of the orbital
  - c. Indicates relative distance from the nucleus
  - d. The higher the value, the higher the energy level
  - e. Positive integer.
2. The **angular quantum number**:
  - a. Use the symbol ( $l$ )
  - b. An interger value from  $0 \rightarrow n-1$
  - c. Indicates the number of nodes
  - d. Possible  $l$  values =  $n$
3. The **magnetic quantum number ( $m_l$ )**
  - a. An integer from  $-l$  through 0 to  $+l$
  - b. Number of values is  $2l+1$
  - c. Total number of  $m_l$  values for a given value of  $n$  is  $n^2$

What is the difference between an angular node and a radial node? (see below)

<sup>1</sup> [https://en.wikipedia.org/wiki/Quantum\\_harmonic\\_oscillator#/media/File:HarmOsziFunktionen.png](https://en.wikipedia.org/wiki/Quantum_harmonic_oscillator#/media/File:HarmOsziFunktionen.png)



Fill in the greyed sections with your answers.

Principal quantum number (n) or energy level (shell)	Total number of subshells in n	Possible values of l, this is the angular quantum no.	Possible values of m <sub>l</sub> , this is the magnetic quantum no.	Subshells for given n	Number of orbitals	Shape of subshell	Max # e. in each subshell	Max # total e in a <b>shell</b>
1	1	0	0	s	1	sphere	2	2
2	2	0	0	s	1	sphere	2	8
		1	$\pm 1, 0$	p	3	dumbbell	6	
3								
4								
5								
					g			
6								
					g			
					h			

Answer the following questions that relate to quantum numbers.

- Looking at the handouts, we can conclude that numbers are used to identify  $l$  and letters are used to identify  $m_l$ .
- An atomic orbital is thought of as a region of space in which there is a  $\psi^2$  probability of finding an electron.

- Electrons are organized into  $s$ - and  $p$ - which identified where they spend 90% of their time.
- Is  $n = 12$  a valid principle quantum number? Explain.
- The quantum numbers listed below are for 4 different electrons in the same atom. List them by order of increasing energy.

Quantum numbers	Orbital designation		List energy from low to high
$n=4, l=0, m_l=0, m_s=1/2$			
$n=3, l=2, m_l=1, m_s=1/2$			
$n=3, l=2, m_l=-1, m_s=1/2$			
$n=3, l=1, m_l=1, m_s=1/2$			

Do any have the same energy?

Which ones?

- What is the subshell designation ( $2p$ , for example) for the following values?
  - $n=4, l=3$
  - $n=5, l=1$
  - $n=3, l=-0$
  - $n=6, l=3$
  - $n=9, l=2$
- What are the maximum number of electrons expressed by the following quantum numbers.
  - $n=3$ ,
  - $n=4, l=2$ ,
  - $n=3, l=1, m_l=-1$
  - $n=7, l=4, m_l=3, m_s=1/2$
  - $n=5, l=6, m_l=0$
- Describe the general shape of the orbitals that have the following sets of quantum numbers:  $[n, l, m_l]$ 
  - $3, 0, 1$
  - $2, 1, 1$
  - $3, 2, 2$

## 6.6 & 6.7 Representations of Orbitals & Many Electron Systems

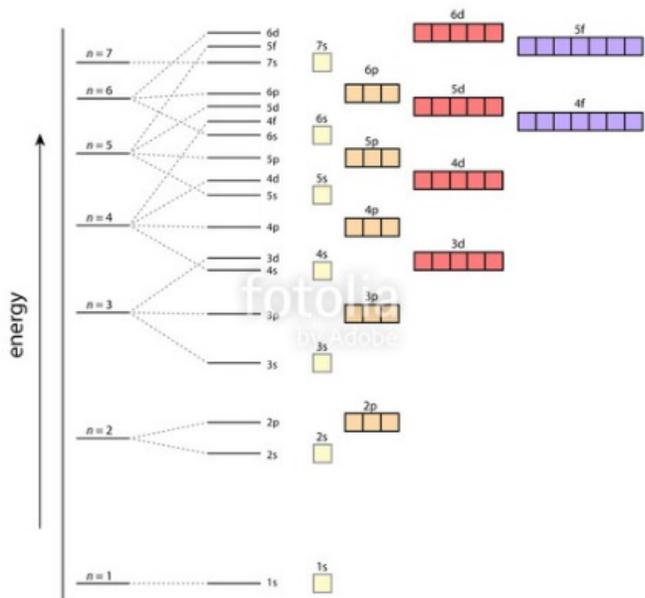
1. Calculate the number of radial nodes in ( $n-l-1$ ):
  - a. 2s orbital
  - b. 9dxy orbital
  - c.  $5f_y^3$
2. Calculate the number of radial nodes in: ( $l$ )
  - a. 2s orbital
  - b. 9dxy orbital
  - c.  $5f_y^3$
3. An atomic orbital is thought of as a region of space in which there is a \_\_\_\_\_ probability of finding an \_\_\_\_\_.
4. Electrons are organized into \_\_\_\_\_ and \_\_\_\_\_ which are used to identify where they spend 90% of their time.
5. Two notations are to represent electrons. When the electron is expressed as a quantum number,  $m_l$  is either \_\_\_\_\_ or \_\_\_\_\_, When the electron is noted as a particle in an electron configuration, we use a half arrow pointed either up for + \_\_\_\_\_ or down for - \_\_\_\_\_. This satisfies the Pauli exclusion Principle.
6. The Pauli Exclusion Principle states that no \_\_\_\_\_ electrons in an atom can have the same \_\_\_\_\_.
7. In a many electron atom, for a given value of  $n$ , the \_\_\_\_\_ increases with the increasing value of  $l$ , the \_\_\_\_\_ momentum quantum number.
8. Orbitals with the same \_\_\_\_\_ are said to be \_\_\_\_\_.

## 6.8 Electron Configurations

1. Hund's Rule of Multiplicity states that when filling \_\_\_\_\_ orbitals, the lowest energy is achieved when the same spin is maximized. Hund's rule maximizes spin by minimizing repulsion.
2. Aufbau is not discussed in the book, but is an important concept because it directs how electrons fill an energy chart in the ground state. In general, Aufbau is simple stated with two rules to help predict electron configurations.
  - Electrons are assigned to orbitals with increasing value of  $n+l$ .
  - For degenerate subshells, electrons are assigned to the subshell with the lower  $n$ .
  - It follows Madelung's rule in which the orbitals of lower  $n+l$  are filled before higher  $n+l$ . The order is 1s2s2p3s3p3d4s4p4d4f and so on. For example, palladium (not one you need to know) the predicted configuration is [Kr]  $5s^24d^9$ , but experimental it's been shown to be [Kr]  $5s^24d^{10}$ . ( $n+l = 5+0$  vs  $4+2$ ).

The rule is based on the total number of nodes because nodes relates to energy. It is a prediction, not an absolute. This is reflected in the general ordering of orbitals for a many-electron atom. Transition metals often do not follow Aufbau, and fill lower d orbitals first. This

explains experimental values based on ionization energies and PES studies. There are two elements that you need to know for the exam which violate Aufbau and Madelung's rules. What are they?



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Subshell electron capacity			
2	6	10	14
		6d	5f
7s	6p	5d	4f
6s	5p	4d	
5s	4p	3d	
4s	3p		
3s	2p		
2s			
1s			

**writing electron configurations:**

Electrons fill from the lowest energy distribution to the highest. Atoms in the ground state, show a characteristic pattern that follows the periodic table as we will see in the next section.

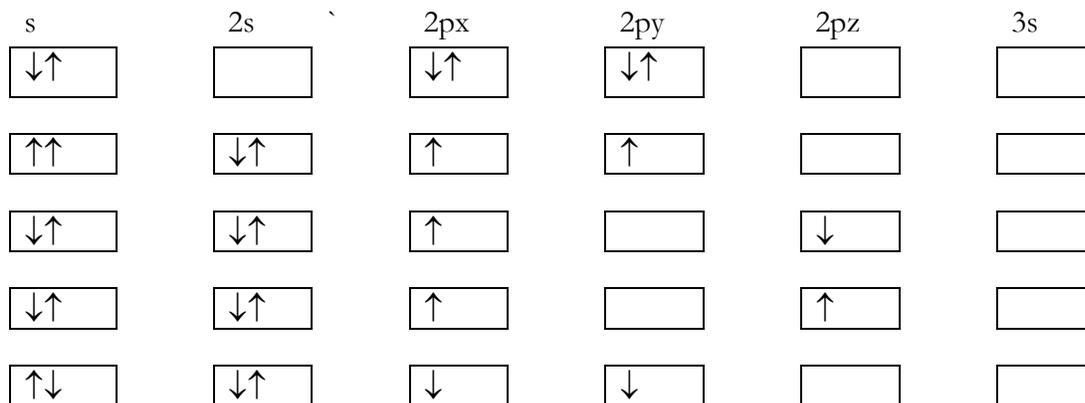
Each line represents an energy level and an orbital.

For example 2s means the electrons are in the 2<sup>nd</sup> period of the periodic table and are in the s block.

When we notate the number of electrons in the block, we use a superscript. This pattern follows

until we reach the p block and degenerate orbitals. Filling in the box diagram shows the distribution of the electrons and must obey Pauli's Exclusion principle and Hund's Rule of Multiplicity.

Identify the excited state(s), the ground state, and the forbidden state(s) using Aufbau, the Pauli Exclusion Principle, and Hund's Rule for the box diagram of a carbon atom.



## 6.9 Electron Configurations and the Periodic Table

Electron configurations represent the lowest energy distribution of electrons in the energy levels of the ground state of the atom. They do not give us the actual values. The recurring pattern in electron configurations is the bases for recurring patterns in chemical behavior.

- The periodic table is a good guide for configurations.
- Periods have values n
  - The 's' and 'p' blocks rows = n
  - The 'd' block, n-1
  - The 'f' block n-2

There are several accepted methods for writing electron configurations. Each has merit, and should be mastered.

1. Full electron configuration following the periodic table:
  - a. You are given an periodic table, counting electrons, blocks, and rows, makes filling the electron configuration easy.

1s→2s→2p→3s→3p→4s→3d→4p→5s→4d→5p→6s→4f→5d→6p→7s→5f→6d

${}_{81}\text{Ti}: 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^1$

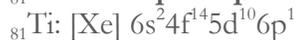
2. Energy order:
  - a. Once the electron configuration is written in periodic table order, sometimes it is beneficial to reorganize the fill by energy levels, n, and l.
  - b. Electrons are lost and gained in the highest **nl** subshell.

- c. Writing in energy order helps this process (I am showing ONLY the orbitals from the periodic table, 5g, 6f, 6g, 6h, 7d etc are not on the table but are theoretical quantum values.

$1s \rightarrow 2s \rightarrow 2p \rightarrow 3s \rightarrow 3p \rightarrow 3d \rightarrow 4s \rightarrow 4p \rightarrow 4d \rightarrow 4f \rightarrow 5s \rightarrow 5p \rightarrow 5d \rightarrow 5f \rightarrow 6s \rightarrow 6p \rightarrow 6d \rightarrow 7s$



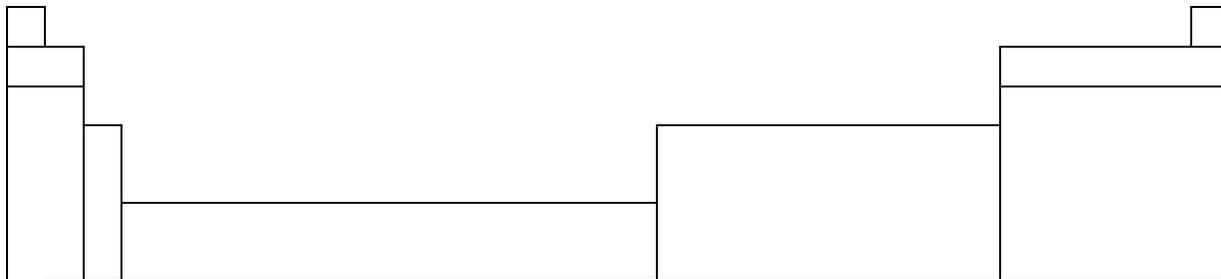
3. Condensed or noble gas configurations:
- Its short! Its based on Aufbau.
  - The noble gas core replaces most of the configuration.



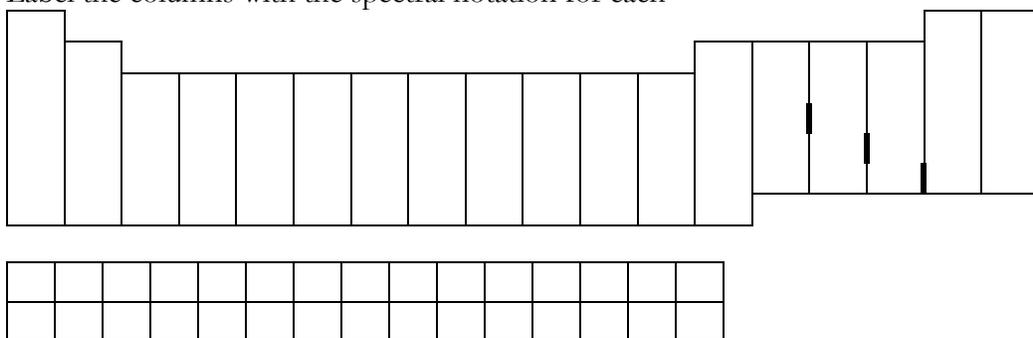
Fill in the table with the appropriate information for the following elements

Symbol	Electron configuration	Noble gas configuration
Rb		
Ba		
Ti		
Mo		
Si		
Pb		
Cl		

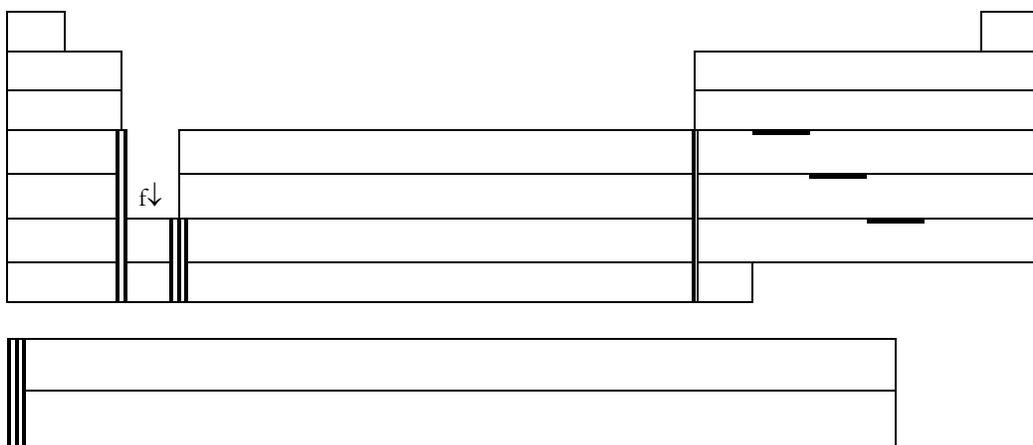
Fill in the blocks



Label the columns with the spectral notation for each



Fill in the periodic table below with the appropriate periods and sublevels



General fill for table based on families. Fill in the missing values.

Group/families	General configuration of last electron in
Alkali Metals (1A)	$ns^1$
Alkaline Earth Metals (2A)	
Group 3A	
Group 4A	
Group 5A	
Group 6A (Chalagons)	
Halogens (7A)	
Transition elements	$(n-1) d^{(column\#)}$
Inner transition elements	

Fill in the table with the appropriate information

Symbol	Block of last e in	Period of last e in	Notation of last fill sublevel	Group No.	No. Valence e
Rb					

Symbol	Block of last e in	Period of last e in	Notation of last fill sublevel	Group No.	No. Valence e
Ba					
Ti					
Mo					
Si					
Pb					
Cl					

For the exam:

You should be able to write the electron configurations for the first 36 elements and the representative elements.