Chapter 5 Atomic Structure and Light

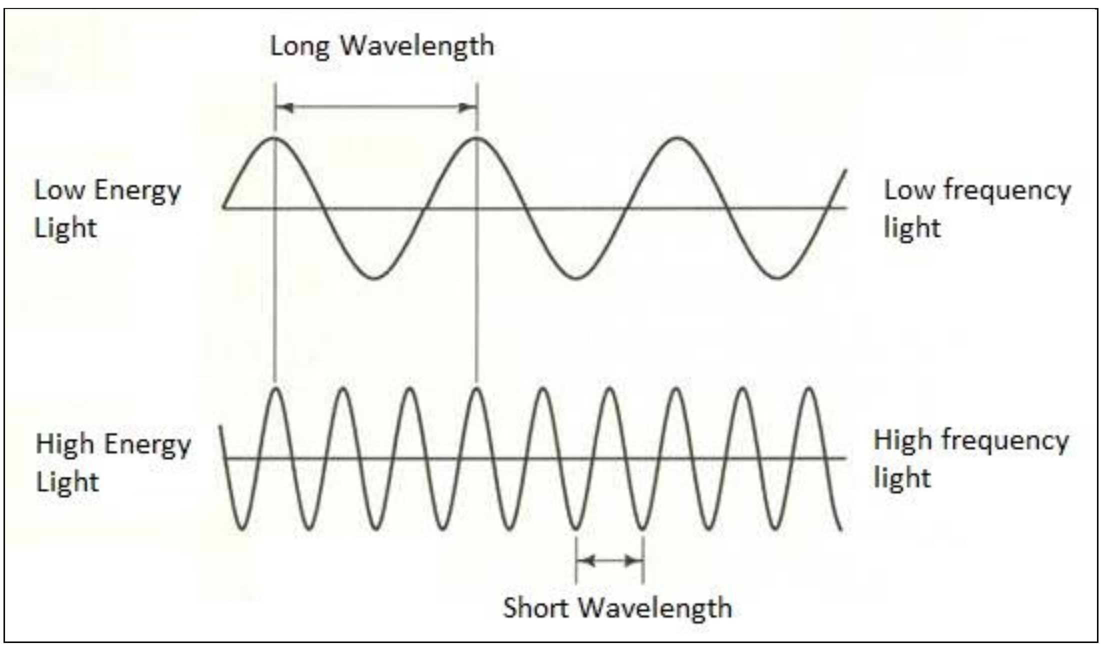
5.1 The Electromagnetic Spectrum

The Electromagnetic Spectrum is the range of all possible frequencies of light. First we must understand light which travels in waves.

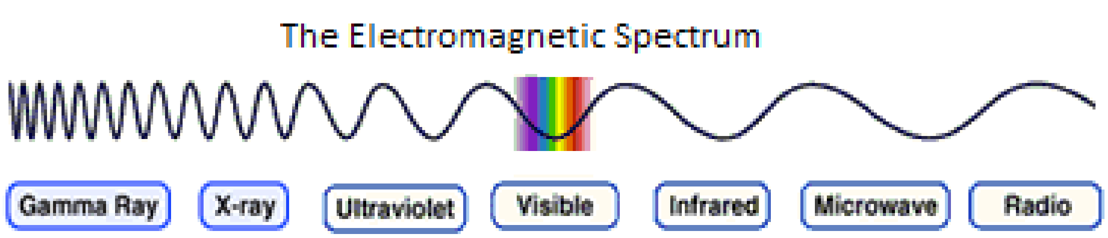
Wavelength (λ) is the distance between peaks on adjacent waves.

Frequency (ν) is the number of wave cycles completed in one second. (like a beat)

⇒ As wavelength ↑, the frequency ↓, and the energy ↓
⇒ As wavelength ↓, the frequency ↑, and the energy ↑



The electromagnetic spectrum, below, is a continuous spectrum from high gamma energy to low radio energy



You should know these in order: (high E) Gamma, X ray, UV, Visible, IR, Microwave, TV, Radio (low E)

Light that we can observe with the naked eye falls within the visible region of the spectrum.

⇒ When white light is passed through a prism, it separates into a continuous spectrum of all wavelengths of visible light. (ROYGBIV)

⇒ Range of visible spectrum is from violet (400 nm) to red (700 nm).

Note that all light travels at light speed! Gamma and radio waves travel at the same speed.

5.2 Bohr's Model of the Atom and Line Spectra

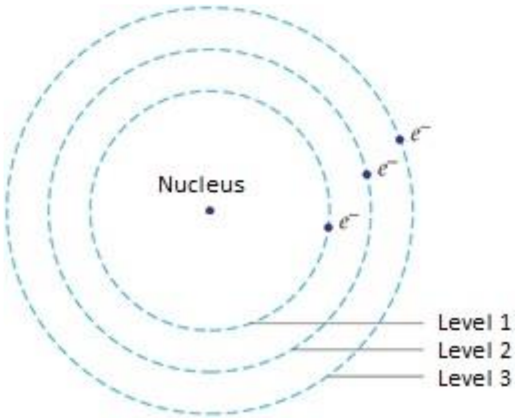
In 1900, Max Planck proposed the controversial idea that energy was emitted in small bundles called quanta. (Energy is not continuously emitted). He defined an individual unit of light energy as a **photon**. Analogy: A ball loses potential energy in quantized amounts when it bounces down a stairway whereas a ball loses potential energy continuously if it rolls down a ramp.

Thus stairs are **quantized** but a ramp is continuous. Can you think of other things that are quantized?

In 1913, Neils Bohr proposed that electrons orbit rapidly around the nucleus, occupying circular orbits with distinct **energy levels**. This would mean that electrons are quantized.

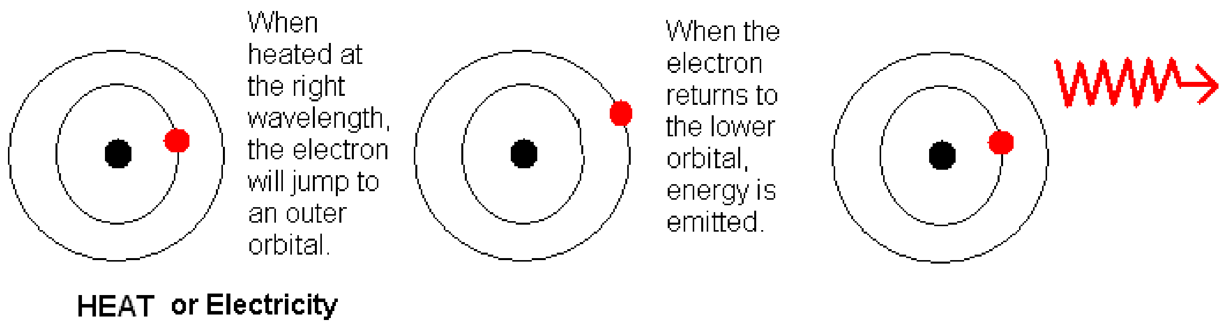
Bohr's model of the atom (1913)

- The electrons orbit around the nucleus like planets orbit around the sun but in 3D space.
- These orbits are called **energy levels**.
- Each orbit has a specific radius and a specific energy.
  - ⇒ The orbit closest to the nucleus is lowest in energy; the energy of the orbit increases with distance from the nucleus.

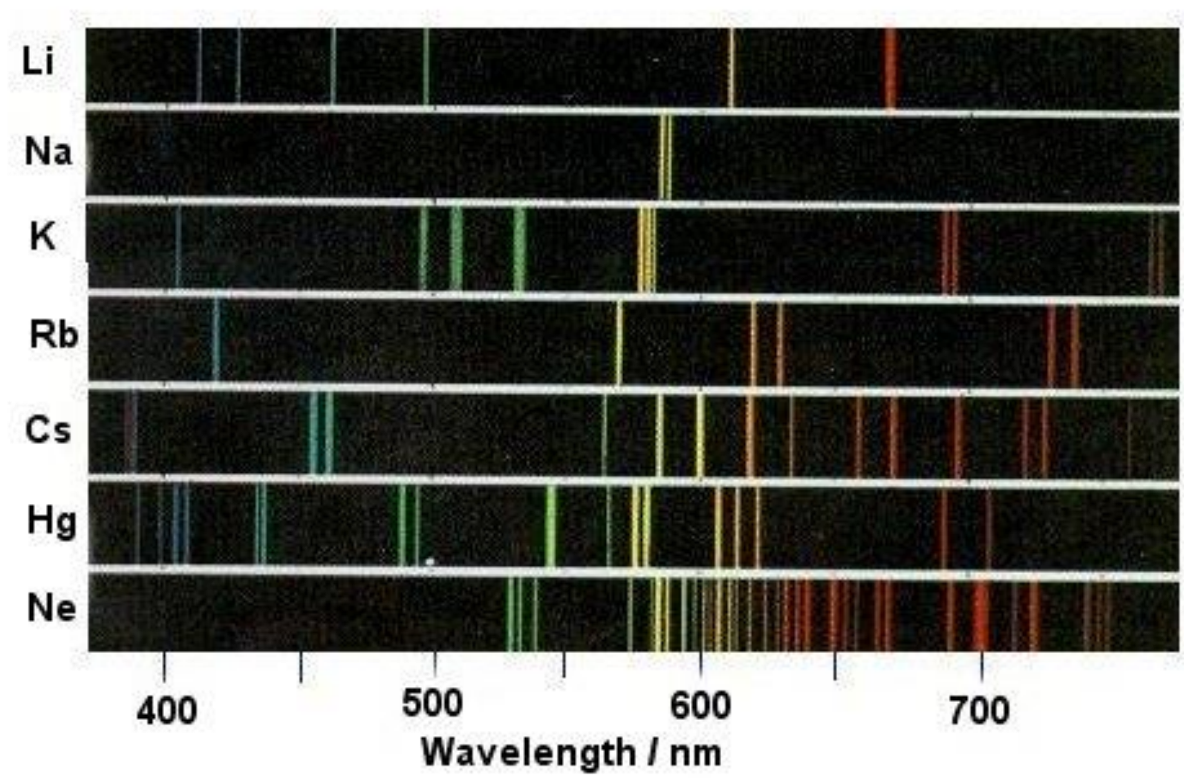


Why did Bohr think the electrons orbit in quantized energy levels?

It is because of the behavior of a gas sealed in a glass discharge tube energized by electricity. When such a gas bulb is plugged in, it gives off light. When the light from the energized gas passes through a prism, an **emission line spectrum** with distinct bands of color is observed. The colors correspond to the wavelengths of emitted light from the **g**as. The energy (or heat) excites the gas allowing the electrons to jump to higher energy levels. Then the "excited" electrons relax back down to their original energy level they release energy as a wave which may be in the visible region so we can see the light. This is how Neon signs work.



Each atom has its own unique line spectrum – this is virtually an **"atomic fingerprint"** that can be used identify the element. Here are several line spectra:



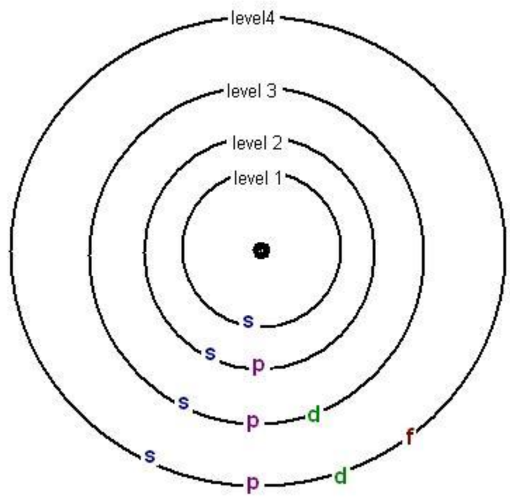
5.3 Electron Levels, Sublevels, and Orbitals

Bohr's model could not explain the emission spectra for all the elements. It later emerged that each **energy level** (numbered 1, 2, 3,…) could be divided into **energy sublevels** named s, p, d, and f.

Each level (n) has n sublevels. The sublevels for the first four energy levels are provided below:

n = 1	one sublevel	1s
n = 2	two sublevels	2s 2p
n = 3	three sublevels	3s 3p 3d
n = 4	four sublevels	4s 4p 4d 4f

In reality, the electron does not move in a fixed orbit. Instead, an electron has a high probability of being found within a given volume or region in space. We call this space an orbital.

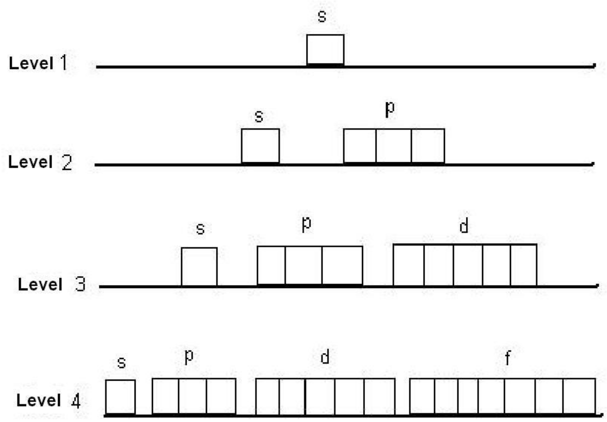


Orbital: region in space where there is a high probability of finding an electron.

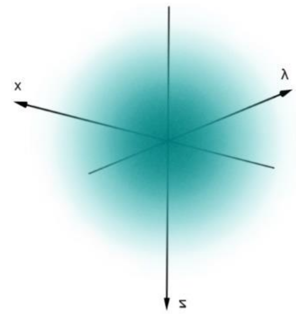
Each orbital can hold a maximum of 2 electrons. Each s sublevel contains 1 orbital and thusntain 2 electrons maximum. Each p sublevel contains 3 orbitals and thus 6 electrons maximum. Each d sublevel contains 5 orbitals and thus 10 electrons maximum. And each f sublevel contains 7 orbitals and thus 14 electrons maximum.

Sublevel	**# of orbitals**	**Max # of e^- in sublevel**
s	1	2
p	3	6
d	5	10
f	7	14

So orbitals exist in the sublevels which are on the levels. Thus we get this representation for the first four levels where a box represents an orbital and can contain 2 electrons.



Sizes and Shapes of Orbitals

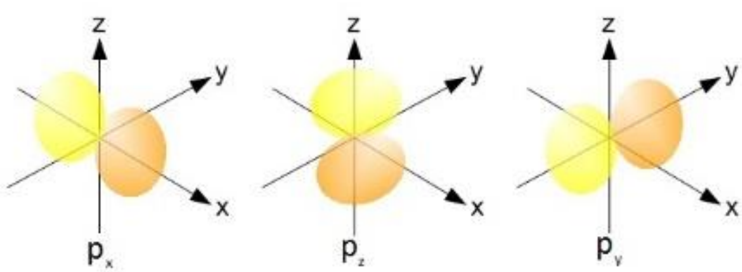


S orbitals are spherical

⇒ As the level number (n)↑, the size and energy of the orbitals ↑

⇒ The blue color indicates the probability of finding an electron in that space.

P orbitals resemble dumbbells, lying along the x,y, and z axis.



Note: You do not need to know about the d or f orbitals.

5.4 Electron Configurations

Electron Configuration: Shorthand description of the arrangement of electrons by sublevel.

Sublevels are filled in order of increasing energy, not physical location.

1s < 2s < 2p < 3s < 3p < 4s

⇒ You will only need to know these sublevels to write electron configurations for the 1st 20 elements in this class.

Once a sublevel has the maximum number of electrons it can hold, it is considered “filled.” Remaining electrons must then be placed into the next highest energy sublevel, and so on, until you are out of electrons for that atom.

Example: Helium. He has 2 electrons. They can both be put into the 1s sublevel. Thus the electron configuration for He is simply $1s^2$. We use subscripts to indicate how many electrons are in that sublevel.

Example: Carbon. C has 6 electrons. We can put 2 into the 1s sublevel, 2 more into the 2s sublevel, and then the last 2 into the 2p sublevel: $1s^2 2s^2 2p^2$

Note: Be able to write electron configurations for elements #1-20.

The Periodic Table actually corresponds to the order of energy sublevels. The sublevels correspond to blocks of elements on the periodic table.

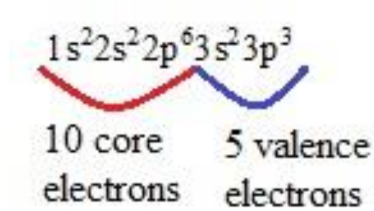
IA																	VIIIA
1 H 1.01	IIA		**s**		**p**		**d**		**f**			IIIA	IVA	VA	VIA	VIIA	2 He 4.00
3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.31	IIIB	IVB	VB	VIB	VIIB	VIIIB	VIIIB	VIIIB	IB	IIB	13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (99)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.75	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57 La* 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.85	75 Re 186.21	76 Os 190.2	77 Ir 192.22	78 Pt 195.09	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	89 Ac* (227)	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (264)	108 Hs (265)	109 Mt (266)	110 Ds (271)	111 Rg (272)	112 Cn (285)	113 Uut (284)	114 Fl (289)	115 Uup (288)	116 Lv (293)	117 Uus (294)	118 Uuo (294)

*Lanthinides	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.4	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
*Actinides	90 Th 232.04	91 Pa (231)	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

Elements in the last column of the Periodic Table are called “Noble gases.” The Noble gases (Group 8A) have completely filled s and p orbitals. This makes them very stable and inert. So having your electron configuration end with s^2p^6 makes you very happy and stable!

Electrons in inner levels are called "**core electrons**" since they are more stable (less reactive) when they belong to levels with full s and p sublevels. The electrons in the outermost unfilled level are the ones that are reactive and form bonds with other atoms. They are called valence electrons.

5.5 Valence Electrons



Consider the element phosphorus, P, which has a configuration of $1s^22s^22p^63s^23p^3$ because it has 15 electrons total. The electrons in the first and second levels are core, while the electrons in the third level (outermost unfilled) are valence.

Valence electrons form chemical bonds between atoms and dictate the properties and chemical behavior of an element.

For Group A elements: Group # = # of valence electrons

Group	IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA
# valence e^-	1	2	3	4	5	6	7	8

The exception is He which although in group VIIIA has only 2 valence electrons because it only has two electrons total.

5.6 Ion Formation and Ion Electron Configurations

This section is very important – please read carefully! So why do atoms make compounds? Why doesn't the world simply consist of atoms? Why do they bond together? Well thank goodness atoms do bond otherwise we would not be here! The reason is because all the atoms want to have a noble gas electron configuration ending in s^2p^6 which means 8 valence electrons otherwise known as the octet rule. So atoms gain or lose electrons to form ions (atom with a charge) so they can achieve a noble gas electron configuration. This makes the atoms, now ions, happy and stable! Noble gases are very happy and stable with their full s and p sublevels, s^2p^6. Of course He just has a full s sublevel, s^2.

Example: $Na \xrightarrow{\text{loses 1 electron}} Na^+$

Electron configuration: $1s^22s^22p^63s^1$ becomes $1s^22s^22p^6$ (happy and stable like Ne)

Note that the sodium atom has one valence electron, the $3s^1$ electron, and that by losing it Na then has the electron configuration of neon with ten electrons. Sodium has lost one negative electron therefor it now has a positive 1 charge. So Na^+ is an ion and is much more stable and happy than Na the atom.

Example: $S \xrightarrow{\text{gains 2 electrons}} S^{2-}$

Electron configuration: $1s^22s^22p^63s^23p^4$ becomes $1s^22s^22p^63s^23p^6$ (stable like Ar)

Note that the sulfur atom has 6 valence electrons, $3s^23p^4$, and that by gaining two the S atom then has the electron configuration of argon with 18 electrons total. Sulfur has gained two negative electrons therefore it now has a negative 2 charge. So S^{2-} is an ion and is much more stable and happy than S the atom.

5.7 Ionic Charges and Isoelectronic

In general:

Metals lose electrons from their valence shell and become **positively charged ions = cations**

Group	**Group IA metals**	**Group IIA metals**	**Group IIIA metals**
Charge	**+1**	**+2**	**+3**
Example	**Li^+**	**Mg^{2+}**	**Al^{3+}**

Nonmetals gain electrons, adding electrons to their valence shell and become **negatively charged ions = anions**

Group	**Group VA nonmetals**	**Group VIA nonmetals**	**Group VIIA nonmetals**
Charge	**–3**	**–2**	**–1**
Example	**N^{3-}**	**O^{2-}**	**F^-**

- Charges on ions can be shown with the sign before or after the number—e.g. magnesium ion can be shown as Mg^{+2} or Mg^{2+}.
- The 1 is omitted for ions with a +1 or –1 charge : Na^+ or F^-
- The IVA column can be charged +4 or -4, and Sn and Pb can even be +2 charged, so this column is not known as the charge can vary.

A + charge means electrons have been lost, so we subtract the number of electrons given by its ionic charge, there are less electrons than the neutral atom

A – charge means electrons have been gained, so we add the number of electrons given by its ionic charge, there are more electrons than the neutral atom

In the examples from the previous section, Na^+ and Ne have the same electron configuration which makes these two species isoelectronic! S^{2-} and Ar are also isoelectronic.

Isoelectronic: two or more species that have the same number of electrons and the same electron configuration.

5.7 Protons, Neutrons and Electrons in Ions

Ions have the SAME number of protons as their original atom, the number of protons never changes in a chemical reaction. Similarly the number of neutrons does not change either. However as electrons are gained or lost, the number of electrons then does change from the original atom. Consider the table below:

	^{23}Na atom	$^{23}Na^+$ ion	^{34}S atom	$^{34}S^{2-}$ ion
number of p^+	11	11	16	16
number of n	12	12	18	18
number of e^-	11	10	16	18

You should be able to fill in a similar table on your own. Remember the neutrons are found by subtracting the number of protons from the mass number.

Practice Problems

1. How many electrons can fit into a d subshell?
2. How many electrons can fit into the 3rd level?
3. How many electrons can fit into an orbital?
4. How many sublevels are on the fourth level?
5. How many orbitals are on the 2nd level?
6. How many orbitals are in a d sublevel?
7. Write electron configurations for Ne, Ar, K, B, and S atoms.
8. How many valence electrons are in Be, Ca, Si, Cl, and O atoms?
9. What type of radiation is between X rays and visible light?
10. What type of radiation has the lowest frequency?
11. As frequency increases what happens to wavelength and energy?
12. Briefly explain how a neon sign works.
13. Why does a sodium atom lose one electron?

14. Why does a fluorine atom gain one electron?
15. What are the electron configurations for Mg^{2+}, P^{3-}, and O^{2-} ions? With what atom are they isoelectronic?
16. What charge will the following atoms have once they gain/lose electrons to become stable ions? Ba, I, Al, S, Br, Li.
17. Fill in the following table:

	^{81}Br	$^{81}Br^{-}$	^{41}Ca	$^{41}Ca^{2+}$
number of p^+				
number of n				
number of e^-				

Answers to Practice Problems

1. ten
2. eighteen
3. two
4. four: 4s, 4p, 4d, and 4f are the sublevels on level four.
5. four: one s orbital and three p orbitals
6. five
7. Ne $1s^22s^22p^6$ Ar $1s^22s^22p^63s^23p^6$ S $1s^22s^22p^63s^23p^4$
   K $1s^22s^22p^63s^23p^64s^1$ B $1s^22s^22p^1$
8. Be – 2 val e^-, Ca – 2 val e^-, Si - 4 val e^-, Cl - 7 val e^-, and O - 6 val e^-
9. UV, ultraviolet light
10. Radio waves
11. As frequency increases, energy increases but wavelength decreases.
12. A typical neon sign works by having neon gas in a glass tube that can be plugged in giving energy to the gas. The electrons in the gas atoms receive that energy and "jump" to a higher level as they absorb that energy. Then the electrons release the energy and jump back down to a lower level releasing that energy as a wave that can be in the visible region thus creating light that our eyes detect.
13. A sodium atom has only one outer valence electron, the $3s^1$ electron. If it loses that one electron it becomes like Ne or $1s^22s^22p^6$ which is very stable so that makes Na^{+1} happy. Having 8 outer valence electrons makes an atom stable and happy.
14. Flourine has 7 valence electrons, $1s^22s^22p^5$. It the atom gains one electron then it becomes like Ne or $1s^22s^22p^6$ which is very stable so that makes F^{-1} happy.

15. Mg^{2+} is $1s^2 2s^2 2p^6$ and is isoelectronic with Ne

P^{3-} is $1s^2 2s^2 2p^6 3s^2 3p^6$ and is isoelectronic with Ar

O^{2-} is $1s^2 2s^2 2p^6$ and is isoelectronic with Ne

16. Ba^{2+}, I^{1-}, Al^{3+}, S^{2-}, Br^{1-}, Li^{1+}.

17. Fill in the following table:

	^{81}Br	$^{81}Br^{-}$	^{41}Ca	$^{41}Ca^{2+}$
number of p^+	35	35	20	20
number of n	46	46	21	21
number of e^-	35	36	20	18