V. Atomic Structure Answers

0. As far as we are concerned in this course, 99.99% of the matter in the universe is composed of three subatomic particles. What are the names and charges of these particles? Where in the atom do these particles exist?

proton: symbol is p^+ and charge is +1	Exist in nucleus (small space at the center of the atom)
neutron: symbol is n and charge is 0	Exist in nucleus (small space at the center of the atom)
electron: symbol is e^- and charge is −1	Move rapidly in orbitals around the outside of atom

1. Given the elemental symbol X shown below, what do Z, M, C, and H stand for?

Z is called the atomic number. Z = the number of p^+.
M is called the mass number. M = the number of p^+ plus the number of n.
C is the net charge. C = the sum of the charge of the p^+ plus the charge of the e^-.
H is the number of atoms of this element present in a molecule or compound ion. (Ex.: One carbon atom and two oxygen atoms are present in CO_2.)

M C
Z **X** H

2. How many p^+, n, and e^- are present in each of the following, and would you expect the indicated material to be stable? (See Supplemental Information part of the web page for a Periodic Table.)

	p^+	n	e^-	stable? (answers re. only e^- stability)
^{6}Li	3	3	3	not nearly as stable as Li^+
$^{81}Br^-$	35	46	36	Yes, much more so than Br^0
^{196}Pt	78	118	78	Yes, *but you might not know this*.
^{41}Ar	18	23	18	Yes! It's an inert gas!
$^{90}Sr^{2+}$	38	52	36	Yes, much more so than Sr^0

3. Describe the number, shape, and relative size of:

a) 4*s* atomic orbitals — There is one *s* orbital per principle quantum number. (Ex.: only one 1*s* orbital, only one 2*s* orbital, *etc*.) *s* orbitals have one spherical lobe per orbital. 4*s* orbitals have a larger radius than 3*s* orbitals, but a smaller radius than 5*s* orbitals.

b) 3*p* atomic orbitals There are three *p* orbitals per principle quantum number. (Ex.: Three 2*p* orbitals, *etc*.) *p* orbitals have two compressed drop shaped lobes per orbital. 3*p* orbitals are larger than 2*p* orbitals, *etc*..

4. Define the term *valence electron*.

The e^- occupying the outermost *s* and *p* orbitals are referred to as valence e^-. These are the e^- that are lossed or gained during ion formation or shared during covalent bond formation.

5. Write electron configurations for and circle the valence electrons in:

a) N $1s^2 2s^{(2)} 2p^{(3)}$

b) Kr $1s^2 2s^2 2p^6 3s^2 3p^6 4s^{(2)} 3d^{10} 4p^{(6)}$

c) Cl^- $1s^2 2s^2 2p^6 3s^{(2)} 3p^{(6)}$

d) Ca^{2+} $1s^2 2s^2 2p^6 3s^{(2)} 3p^{(6)}$

6. Comment on the electronic stability of 5a)-d).

a) N Nitrogen atoms are quite unstable because they do not have a full valence shell.

b) Kr Krypton atoms are quite stable because they do have a full valence shell.

c) Cl^- Chloride ions are quite stable because they do have a full valence shell.

d) Ca^{2+} Calcium ions are quite stable because they do have a full valence shell.