Name___________________________

Section___________ Date_______

Electronic Configuration Worksheet

Scientists have learned from experiment that electrons *in atoms* occupy specific energy levels. These energy levels form an orderly pattern, and once this pattern is understood, it is rather simple to figure out, for any element, where the electrons are, i.e., which energy levels they occupy. Once we know this electron arrangement (also called the electronic configuration), we can predict a great deal about how these electrons will behave in chemical reactions. And this is the basis for chemistry--a fundamental understanding of the nature of atoms so that we can begin to understand and predict what they do.

Look at the diagram on the next page. It shows major, or principal energy levels. These are numbered consecutively 1, 2, 3, ...; the first 7 are shown on the diagram. It also shows that each of these levels is split into sublevels. The first principal energy level consists of a single sublevel; the second one has two sublevels, and so on. For levels 5, 6 and 7, not all the sublevels are shown. The sublevels are designated by letters: s, p, d, f, g, h, i, ... (the reason for choosing these letters is historical.)

Notice that as energy increases, the levels become more and more closely spaced. Although the exact spacing changes from element to element, the same general pattern exists for all elements--a large gap between levels one and two, a smaller gap between levels two and three, and so on.

Exercise:

Complete the list of sublevels For principal energy level 4:

4_s_, 4___, 4___, 4___.

List the sublevels for principal energy level 6: ___, ___, ___, ___, ___, ___

What sort of experiment indicates the existence of discrete energy levels?

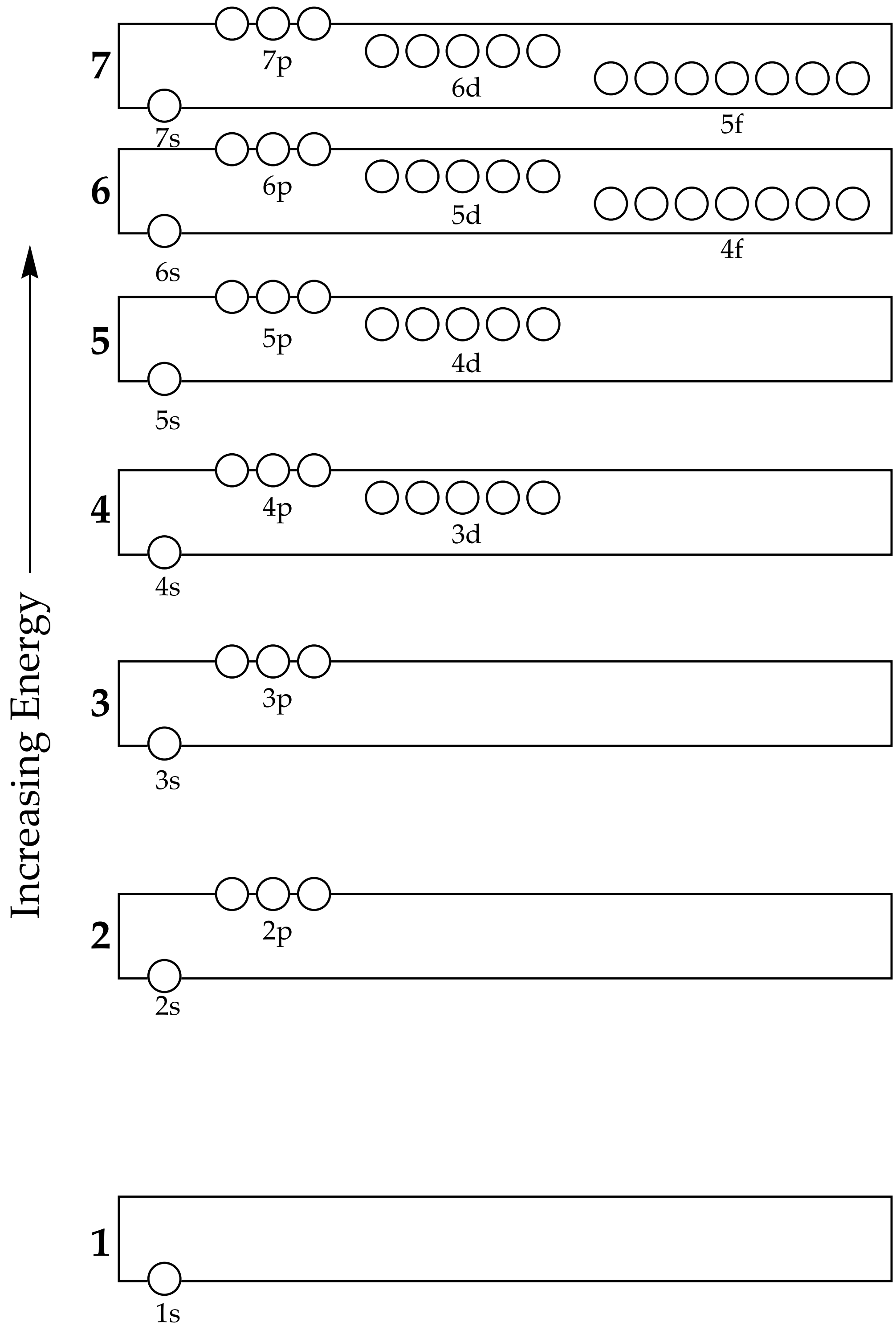
The energy levels and sublevels by themselves do not tell the whole story. Because the spacing between energy levels becomes smaller and smaller as energy increases, the sublevels of different principal energy levels overlap each other. Note on the diagram on the next page that the 3d sublevel, for instance, is higher in energy than the 4s.

One way to learn the order of the sublevels is simply to remember the arrangement on the diagram on the next page. Another is to remember the arrangement shown in Figure 11-14 on p. 467.

Exercise:

Copy the diagram from Figure 11-14
on p. 467 in the space at the right.

Electronic Energy Levels



No orbitals higher in energy than the 7p are needed to describe the unexcited electrons in every known element (as of 2003).

The next step is to find out how many electrons can fit into each sublevel. The answer to this question comes from theory, but is verified by experiment. Sublevels are divided into orbitals. An *orbital* is not the same as an *orbit*. An orbit is a well-defined path, like the earth's orbit around the sun. An orbital is a mathematically defined volume of space in an atom, molecule or ion, within which an electron can be found some specified percentage of the time. Electrons in orbitals in a given sublevel have the same energy.

Exercise:

Fill in the answers (number and diagram) for f, g, and h sublevels, below:

s sublevels have 1 orbital __

p sublevels have 3 orbitals __ __ __

d sublevels have 5 orbitals __ __ __ __ __

f sublevels have ___ orbitals

g sublevels have ___ orbitals

h sublevels have ___ orbitals

Experiments have shown us that <u>*each orbital may hold a maximum of 2 electrons*</u>. When two electrons occupy a single orbital, their spins are aligned in opposite directions: <u>↑↓</u>

Thus an orbital may have no electrons (empty): __; one electron: <u>↑</u>, or two electrons: <u>↑↓</u>.

Exercise:

For each of the following sublevels, indicate a diagram for the maximum number of electrons:

<u>orbital</u>		<u>how many electrons?</u>
2p	<u>↑↓</u> <u>↑↓</u> <u>↑↓</u>	<u>6</u> electrons
2s		___ electrons
4p		___ electrons
3d		___ electrons
5f		___ electrons
6s		___ electrons
5g		___ electrons
6f		___ electrons

What is the maximum number of electrons in principal energy level 1 _________, principal energy level 2 _________, principal energy level 3 _________?

We have now looked at the basic rules for energy levels, sublevels, and orbitals. Now we will use this information to find the arrangement of electrons of different elements. These arrangements are called *orbital diagrams*.

Some guidelines:

1. The energy sequence of the sublevels is the same for all elements.
2. Electrons go into the lowest energy orbital that has room available. This is the "Pauli Exclusion Principle."
3. The maximum number of electrons in an orbital is 2.

We list the first few elements and their sublevels. Electrons fill orbitals from the bottom up:

	${}_1$H	${}_2$He	${}_3$Li	${}_4$Be	${}_5$B
2p	— — —	— — —	— — —	— — —	$\uparrow$ — —
2s	—	—	$\uparrow$	$\uparrow\downarrow$	$\uparrow\downarrow$
1s	$\uparrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$

The fifth electron in boron (B) could go into any one of the three 2p orbitals. But when we come to carbon (C), we have two distinct choices for the sixth electron:

	This? ${}_6$C	or this? ${}_6$C
2p	$\uparrow$ $\uparrow$ —	$\uparrow\downarrow$ — —
2s	$\uparrow\downarrow$	$\uparrow\downarrow$
1s	$\uparrow\downarrow$	$\uparrow\downarrow$

The answer again comes from experiment, and gives us Guideline #4:

4. Electrons will occupy orbitals of equal energy singly before pairing up. This statement is known as Hund's Rule of Maximum Multiplicity, or, more simply, as "Hund's Rule."

Exercise:

Fill in the electrons of the second row elements from N through Ne:

	${}_6$C	${}_7$N	${}_8$O	${}_9$F	${}_{10}$Ne
2p	— — —	— — —	— — —	— — —	— — —
2s	—	—	—	—	—
1s	—	—	—	—	—

(The 2p electrons in carbon can occupy any two of the 2p orbitals.)

Exercise:

Fill in the electrons for the third-row elements:

	$_{11}Na$	$_{12}Mg$	$_{13}Al$	$_{14}Si$	
3p	— — —	— — —	— — —	— — —	3p
3s	—	—	—	—	3s
2p	— — —	— — —	— — —	— — —	2p
2s	—	—	—	—	2s
1s	—	—	—	—	1s

	$_{15}P$	$_{16}S$	$_{17}Cl$	$_{18}Ar$	
3p	— — —	— — —	— — —	— — —	3p
3s	—	—	—	—	3s
2p	— — —	— — —	— — —	— — —	2p
2s	—	—	—	—	2s
1s	—	—	—	—	1s

In addition to the orbital diagrams the electron arrangement can be written in shorter form called *electron configuration*:

H	He	Li	Be	B	and so on.
$1s^1$	$1s^2$	$1s^22s^1$	$1s^22s^2$	$1s^22s^22p^1$	

Exercise:

Write the short notation electron configurations for elements 6-18:

Beginning with K, the fourth period elements fill in the 4s, 3d and 4p sublevels, in that order. In the orbital diagrams below, the first 18 electrons (in the previously filled sublevels) are omitted:

	$_{19}K$	$_{20}Ca$	$_{21}Sc$	$_{22}Ti$	
3d	— — — — —	— — — — —	↑ — — — —	↑ ↑ — — —	3d
4s	↑	↑↓	↑↓	↑↓	4s

	$_{23}V$	$_{24}Cr$	$_{25}Mn$	
3d	↑ ↑ ↑ — —	↑ ↑ ↑ ↑ ↑	↑ ↑ ↑ ↑ ↑	3d
4s	↑↓	↑	↑↓	4s

Note that Cr is unusual, with configuration $4s^1 3d^5$, and ***not*** $4s^2 3d^4$ as you would expect. The theoretical explanation is that half-filled or fully-filled sublevels are more stable than sublevels that are neither. As the energy levels get closer in energy, starting with the 4s, it takes more energy to pair the electrons in the 4s than it does to promote one to the 3d level. Why is this true? It is true because there is stability gained (energy kick-back) in producing a half-filled or fully-filled 3d sublevel. Therefore an electron will promote from the 4s to the 3d sublevel when the resulting configuration has a half-filled or fully-filled 3d sublevel. All other electron configuration inconsistencies found later in the periodic table can be explained with similar arguments, that the pairing energy is greater than the energy required to promote.

Exercise:

Continue filling in the orbital diagrams below for Fe to Ge. (NOTE: Cu is an exception, similar to Cr.) Under each orbital diagram on this page, write the short notation for the electron configuration.

	$_{26}Fe$	$_{27}Co$	$_{28}Ni$	$_{29}Cu$	
3d	— — — — —	— — — — —	— — — — —	— — — — —	3d
4s	—	—	—	—	4s

	$_{30}Zn$	$_{31}Ga$	$_{32}Ge$	
4p	— — —	— — —	— — —	4p
3d	— — — — —	— — — — —	— — — — —	3d
4s	—	—	—	4s

Now, sharpen your pencil, and in the small space provided on the periodic chart on the next page, write the electronic configuration of the highest filled, or partially filled, energy sublevel for each element up to xenon (element 54).

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

Lanthanides															
Actinides															

Which groups are the representative (main group) elements?

Which sublevels are filling as one goes across a row of representative elements?

Which sublevels are filling as one goes across a row of transition elements?

Which sublevel is being filed by the lanthanides?

Which sublevel is being filled by the actinides?

Which sublevel is being filled by elements $_{104}Rf$ and $_{105}Db$?

Because the inner (noble gas core) electrons of an element are not as important as the outer most electrons, chemists often write electron configurations in an abbreviated form. The noble gas core electrons are represented with the noble gas symbol written in brackets followed by the remaining configuration. *Only the noble gases may be used for core representations*. Write the shorthand electronic configurations for the following:

Si: $[Ne]3s^2 3p^2$ Ag:[Kr] Se:[Ar]

Br:[Ar] Sr:[Kr] Ar: [Ne]

Diamagnetic elements have no unpaired electrons in any orbital; paramagnetic ones have one or more unpaired electrons. Identify which of the following are diamagnetic, and which are paramagnetic (remember that all are one or the other!). Write "dia" or "para":

He: B: C: Ne: Ti: Fe: Zn: S: Sn: