## Answers to Review \#1: Atomic Theory and the Periodic Table (Chapter 3)

1. Know the meanings of, and be able to apply, the following terms:

| atomic number | excited state | isotope | principal quantum number |
| :--- | :--- | :--- | :--- |
| atomic radius | ground state | mass number | quantum number |
| electron configuration | ionization energy | net nuclear attraction $\left(Z_{\text {eff }}\right)$ | shielding effect |
| electronegativity | isoelectronic | orbital | valence electrons |

2. Complete the following table:

| Element | Atomic <br> Number | Number of <br> Protons | Number of <br> Electrons | Number of <br> Neutrons | Charge | Mass <br> Number |
| :---: | :---: | :---: | :---: | :---: | :---: | :---: |
| $\mathrm{Fe}-56$ | 26 | 26 | 23 | 30 | $3+$ | 56 |
| $\mathrm{Co}-58$ | 27 | 27 | 24 | 31 | $3+$ | 58 |
| $\mathrm{Te}-128$ | 52 | 52 | 54 | 76 | $2-$ | 128 |
| $\mathrm{Cd}-112$ | 48 | 48 | 46 | 64 | $2+$ | 112 |

3. Complete the following table:

| Element | Atomic Number | Number of <br> Protons | Number of Electrons | Number of <br> Neutrons | Mass Number | Ground State Electron Configuration of Ion | Three (3) Isoelectronic Ions or Atoms |
| :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: |
| ${ }_{14}^{28} \mathbf{S i}^{4-}$ | 14 | 14 | 18 | 14 | 28 | $[\mathrm{Ne}] 3 s^{2} 3 p^{6}$ | $\begin{aligned} & \mathrm{S}^{2-}, \mathrm{Cl}^{1-}, \mathrm{Ar} \\ & \mathrm{P}^{3-}, \mathrm{K}^{1+}, \\ & \mathrm{Ca}^{2^{2+}}, \mathrm{Sc}^{3+} \end{aligned}$ |
| ${ }_{35}^{80} \mathbf{B r}$ | 35 | 35 | 36 | 45 | 80 | $[A r] 4 s^{2} 3 d^{10} 4 p^{6}$ | $\begin{aligned} & \mathrm{Se}^{2-}, \\ & \mathrm{Rb}^{1+}, \\ & \mathrm{Br}^{1+}, \\ & \mathrm{Kr}, \\ & \mathrm{Sr}^{3+} \end{aligned}$ |
| ${ }_{21}^{45} \mathbf{S c}$ | 21 | 21 | 18 | 24 | 45 | [Ar] | $\begin{aligned} & \mathrm{S}^{2-}, \mathrm{Cl}^{1-}, \mathrm{Ar} \\ & \mathrm{P}^{3-}, \mathrm{K}^{1+}, \\ & \mathrm{Ca}^{2+}, \mathrm{Sc}^{3+} \end{aligned}$ |

4. Be able to interpret electron configurations of neutral atoms:

| Electron Configuration | Group | Period | Element | Common Ion | EN | Atomic Radius | $\begin{aligned} & \mathrm{IE}_{1} \\ & (\mathrm{~V}) \end{aligned}$ | Conducts in pure form? |
| :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: |
| [ He ] $2 s^{2} 2 p^{4}$ | VI | 2 | 0 | $\mathrm{O}^{2-}$ | 3.44 | 0.65 | 13.618 | no |
| [Kr]5s ${ }^{2}$ | II | 5 | Sr | $\mathrm{Sr}^{2+}$ | 0.95 | 2.45 | 5.695 | yes |
| $[X e] 6 s^{2} 4 f^{14} 5 d^{10} 6 p^{5}$ | VII | 6 | At | $A t^{1-}$ | 2.2 | 1.43 | NA | no |
| [ Ne ] $3 s^{2} 3 \mathrm{p}^{2}$ | VI | 3 | Si | $\mathrm{Si}^{4+1-}$ | 1.90 | 1.46 | 8.151 | yes/no (semi) |
| [Xe]6s ${ }^{1}$ | I | 6 | Cs | $\mathrm{Cs}^{1+}$ | 0.79 | 3.34 | 3.891 | yes |

5a) Write the condensed electron configurations for iron, phosphorus, cesium and iodine.

- $\mathrm{Fe}[\mathrm{Ar}] 4 s^{2} 3 d^{6}$
- $P[\mathrm{Ne}] 3 s^{2} 3 p^{3}$
- Cs $[\mathrm{Xe}] 6 \mathrm{~s}^{1}$
- I $[K r] 5 s^{2} 4 d^{10} 5 p^{5}$
$5 b+c)$ Write the predicted condensed electron configurations for chromium, copper, silver and gold. For the elements in question 5b, write the actual (experimental) electron configurations. Explain why the electrons are arranged in this way.
- predicted for Cr is $[\mathrm{Ar}] 4 s^{2} 3 d^{4}$, actual for Cr is $[A r] 4 s^{1} 3 d^{5}$. It is lower energy to have the $d$-orbitals half full, so one of the s-electrons is "promoted" to the d-orbitals so that the d-orbitals are half-full
- predicted for Cu is $[\mathrm{Ar}] 4 s^{2} 3 d^{9}$, actual for Cu is $[\mathrm{Ar}] 4 s^{1} 3 d^{10}$. It is lower energy to have the d-orbitals full, so one of the s-electrons is "promoted" to the $d$-orbitals so that the d-orbitals are full
- predicted for $A g$ is $[K r] 5 s^{2} 4 d^{9}$, actual for $A g$ is $[K r] 5 s^{1} 4 d^{10}$. It is lower energy to have the $d$-orbitals full, so one of the s-electrons is "promoted" to the d-orbitals so that the d-orbitals are full
- predicted for $A u$ is $[X e] 6 s^{2} 4 f^{14} 5 d^{9}$, actual for $A u$ is $[X e] 6 s^{1} 4 f^{14} 5 d^{10}$. It is lower energy to have the $d-$ orbitals full, so one of the s-electrons is "promoted" to the d-orbitals so that the d-orbitals are full

6. Trends on the Periodic Table:
a) know the trends for first ionization energy, electronegativity, atomic radius, ionic radius

- First Ionization Energy (the energy required to remove an electron from a neutral atom in its gaseous state)
highest ionization energy
- Electronegativity (a measure of the attraction that an atom has for the electrons in a bond)

- Atomic Radius (the radius of an atom $=\frac{1}{2}$ the distance between two nuclei of the same type of atom, covalently bonded together)

- Ionic Radius (the radius of an ion)
- the ions of metal atoms are smaller than their parent atoms because they have the same nuclear charge but fewer electrons
- the ions of non-metal atoms are larger than their parent atoms because they have the same nuclear charge but more electrons
- in general, ions with the largest number of protons and smallest number of electrons are the smallest
- in general, ions with the fewest number of protons and largest number of electrons are the largest


## Trends on the Periodic Table (cont.)

6b) be able to explain these trends with reference to the concepts of shielding effect and net nuclear attraction (Zeff)
i) why is an ion of sodium smaller than a neutral sodium atom

- a sodium ion has lost one electron, but the nuclear charge (number of protons) has not changed. A sodium atom has 11 electrons, while a sodium ion has only 10 electrons. Because the ion has fewer electrons to repel one another, the electrons do not spread out as much, and the sodium ion is smaller than its parent atom
ii) why does electronegativity increase across the periodic table
- electronegativity is defined as a measure of the relative attraction that an atom has for the electrons in a bond
- as you move from left to right across a period on the periodic table $(\rightarrow)$ the shielding effect is constant, but the net nuclear attraction (Zeff) increases. This means that the valence electrons experience a stronger overall attraction to the nucleus. As net nuclear attraction increases, electronegativity also increases because the valence electrons that are participating in bonding will be pulled closer and closer to that atom's nucleus
iii) why is an atom of bromine larger than an atom of fluorine
- bromine and fluorine are both in Group VII of the Periodic Table so they both have a net nuclear attraction (Zeff) of $7+$, so net nuclear attraction can not explain the difference in atomic size
- as you move down a group on the periodic table, the shielding effect increases. This means that there are more full electron shells between the nucleus and the valence electrons, and this accounts for the greater size of a bromine atom
iv) why is the first ionization energy of sulfur less than that of chlorine
- first ionization energy measures the amount of energy that is required to remove the first electron from a neutral atom in the gaseous state
- sulfur and chlorine are in the same period so they both have the same shielding effect (there are 10 electrons between the valence electrons and the nucleus). However, sulfur has a 6+ $Z_{\text {eff }}$ (net nuclear attraction) while chlorine has a $7+\mathrm{Z}_{\text {eff }}$. Therefore, chlorine's valence electrons are attracted to the nucleus a little more strongly than sulfur's valence electrons, so the energy required to remove a valence electron from chlorine will be slightly higher

6c) be able to interpret successive ionization energies to identify an element:

| IONIZATION ENERGIES (eV, electron volts) |  |  |  |  |  |  |  | \# of Valence Flortmane | Group \# of Element |
| :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: |
| IE ${ }_{1}$ | $\mathrm{IE}_{2}$ | $\mathrm{IE}_{3}$ | $\mathrm{IE}_{4}$ | $\mathrm{IE}_{5}$ | $\mathrm{IE}_{6}$ | $\mathrm{IE}_{7}$ | $\mathrm{IE}_{8}$ |  |  |
| 9.3 | 18.2 | 153.9 | 217.7 | - | - | - | - | 2 | II |
| 11.3 | 24.4 | 47.9 | 64.5 | 392.0 | 489.8 | - | - | 4 | IV |
| 17.4 | 35.0 | 62.6 | 87.2 | 114.2 | 157.1 | 185.1 | 953.8 | 7 | VII |
| 5.1 | 47.3 | 71.7 | 98.9 | 138.6 | 172.4 | 208.4 | 264.2 | 1 | I |
| 14.5 | 29.6 | 47.4 | 77.5 | 97.9 | 551.9 | 666.8 | - | 5 | V |
| 13.6 | 35.1 | 54.9 | 77.4 | 113.9 | 138.1 | 739.1 | 871.1 | 6 | VI |
| 11.0 | 19.7 | 30.2 | 51.4 | 65.0 | 220.4 | 263.3 | 309.3 | 5 | V |
| 10.4 | 23.4 | 35.0 | 47.3 | 72.5 | 88.0 | 281.0 | 328.8 | 6 | VI |

## 7. Quantum Mechanics:

a) The number of different types of orbitals when $n=5$ is: $\underline{5}$ (there are " $n$ " types of orbitals per quantum level)
b) The number of " $s$ " orbitals in the seventh main energy level $(n=7)$ is: $\underline{1}$ (there is only one " $s$ " orbital per quantum level)
c) The maximum number of electrons that can fit in the third energy level $(n=3)$ is: 18 (the maximum number of electrons that can fit in a quantum level is " $2 n^{2 "}$ )
d) The number of electrons that can be held in the 5-p orbitals $(n=5)$ is: $\underline{6}$ (there are always 3 " $p$ " orbitals, and each orbital can hold a maximum of 2 electrons, for a total of 6 " $p$ " electrons)
e) The number of different types of orbitals when $n=3$ is: $\underline{\mathbf{3}}$ (there are " $n$ " types of orbitals per quantum level)
$f$ ) The number of " $p$ " orbitals in the fourth main energy level $(n=4)$ is: $\underline{3}$ (if there are " $p$ " orbitals in a quantum level, they come in "packages" of 3)
g) The maximum number of electrons that can fit in the fifth energy level $(n=5)$ is: $\underline{50}$ (the maximum number of electrons that can fit in a quantum level is " $2 n^{2 "}$ ")
h) The number of electrons that can be held in the 3 -d orbitals $(n=3)$ is: 10 (there are always 5 " $d$ " orbitals, and each orbital can hold a maximum of 2 electrons, for a total of 10 " $d$ " electrons)
8. Write the quantum numbers that correspond to each of the following electrons:

| Last Electron | $n$ | $\ell$ | $m_{\ell}$ | $m_{s}$ |
| :--- | :---: | :---: | :---: | :---: |
| $4 s^{1}$ | 4 | 0 | 0 | $+\frac{1}{2}$ |
| $3 d^{9}$ | 3 | 2 | +1 | $-\frac{1}{2}$ |
| $2 s^{2}$ | 2 | 0 | 0 | $-\frac{1}{2}$ |
| $5 p^{6}$ | 5 | 1 | +1 | $-\frac{1}{2}$ |
| $1 s^{1}$ | 1 | 0 | 0 | $+\frac{1}{2}$ |
| $3 p^{2}$ | 3 | 1 | 0 | $+\frac{1}{2}$ |

9. Which atom or ion is larger:
a) $\mathrm{Si}, \mathrm{Si}^{4+}, \mathrm{Si}^{4-}$, or $\mathrm{Si}^{2+}$ ? Explain.

All of these species have the same number of protons, but they have different numbers of electrons. The largest species will have the most electrons (because they repel each other and take up more space), so $\mathrm{Si}^{4-}$ is the largest ion.
b) $\mathrm{Sb}^{3+}, \mathrm{Sb}^{5+}, \mathrm{Sb}^{3-}$ or Sb ? Explain.

All of these species have the same number of protons, but they have different numbers of electrons. The largest species will have the most electrons (because they repel each other and take up more space), so $\mathrm{Sb}^{3-}$ is the largest ion.

