**Atomic Structure Answers**

**MULTIPLE CHOICE**

 1. ANS: E

 2. ANS: B

 3. ANS: E

 4. ANS: C

 5. ANS: B

 6. ANS: C

 7. ANS: B

 8. ANS: D

 9. ANS: A

 10. ANS: B

 11. ANS: E

 12. ANS: D

 13. ANS: A

 14. ANS: E

 15. ANS: A

**SHORT ANSWER**

 16. Atomic radius increases as you go down a group of elements in the periodic table. This trend is a result of increasing numbers of electrons occupying increasing numbers of energy levels. The effective nuclear charge changes only slightly and therefore does not offset the increase in size due to the increase in energy levels.

 Atomic radius decreases as you go left to right across a period in the periodic table. The valence electrons are found in orbitals of the same energy level. At the same time, the effective nuclear charge is increasing with the increase in nuclear charge, which results in a greater force of attraction pulling the valence electrons closer to the nucleus. Thus, atomic size decreases.

 17. 1) The principal quantum number, *n*, indicates the energy level of an atomic orbital and its relative size.

2) The orbital-shape quantum number, *l*, indicates the shape of the orbital.

3) The magnetic quantum number, *ml*, indicates the orientation of the orbital.

4) The spin quantum number, *ms*, indicates the direction in which the electron is spinning.

 18. The aufbau principle is the imaginary process of building up the ground state electron structure for each atom, in order of atomic number. When determining the electron configuration of an element, the electrons are written sequentially in orbitals of increasing energy, starting with the electron in the 1*s* orbital.

 19. The Pauli exclusion principle states that no two electrons in an atom have the same four quantum numbers. (In other words, no two electrons can occupy the same orbital with the same spin.) For example, boron’s electron configuration is 1*s*2 2*s*2 2*p*1.

 Hund’s rule states that whenever electrons are added to orbitals of the same energy sub-level, each orbital receives one electron before any pairing occurs. When electrons are added singly to separate orbitals of the same energy sublevel, the electrons must all have the same spin. For example, the electron configuration of nitrogen is 1*s*2 2*s*2 2*px*1 2*py*1 2*pz*1.

 20. a) *ml* = 3, 2, 1, 0 1, 2, 3

b) *ml* = 1, 0, 1

 21. The values of *ml* are 2, 1, 0, 1, 2. The five orbitals described by *l* = 2 are called the *d* orbitals.

 22. 

 23.



3d

4s

|  |  |
| --- | --- |
|  | 3p |
| 3s |  |
|  | 2p |

2s



1s

 24.

|  |  |
| --- | --- |
|  |  |
|  | 3p |
| 3s |  |
|  | 2p |
| 2s |  |
|  |  |
| 1s |  |

**PROBLEM**

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|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Electron | Principle quantum number | Orbital-shape quantum number | Magnetic quantum number | Spin quantum number |
| 1 | 3 | 0 | 0 |  |
| 2 | 3 | 0 | 0 |  |
| 3 | 3 | 1 | -1 |  |
| 4 | 3 | 1 | 0 |  |