Electronic Structure Worksheet 1

Given the following list of atomic and ionic species, find the appropriate match for questions 1-4.

(A) Fe$^{2+}$  (B) Cl$^-$  (C) K$^+$  (D) Cs$^+$  (E) Hg$^+$

1. Has the electron configuration: 1s$^2$ 2s$^2$ 2p$^6$ 3s$^2$ 3p$^6$ 3d$^6$.
2. Has a noble gas electron configuration.
3. Has electrons in f orbitals.
4. Is isoelectronic with gold.

5. Which of the following combinations of particles represents an ion of net charge -1 and of mass number 82?
   (A) 46 neutrons, 35 protons, 36 electrons
   (B) 46 neutrons, 36 protons, 35 electrons
   (C) 46 neutrons, 36 protons, 36 electrons
   (D) 47 neutrons, 35 protons, 35 electrons
   (E) 47 neutrons, 35 protons, 36 electrons

6. One species of element M has an atomic number of 10 and a mass number of 20; one species of element N has an atomic number of 11 and a mass number of 20. Which of the following statements about these two species is true?
   (A) They are isotopes.
   (B) They are isomers.
   (C) They are isoelectronic
   (D) They contain the same number of neutrons in their atoms.
   (E) They contain the same total number of protons plus neutrons in their atoms.

7. Which of the following is not a physical property of a substance?
   (A) density
   (B) solubility
   (C) melting point
   (D) reaction with oxygen
   (E) boiling point

8. Which of the following could be an isotope of chlorine?
   (A) $^{37}$Cl$^{17}$
   (B) $^{17}$Cl$^{17}$
   (C) $^{35}$Cl$^{17}$
   (D) $^{37}$Cl$^{17.5}$
   (E) $^{17}$Cl$^{37}$

9. For a neutral As atom in the ground state, how many electrons have quantum numbers $n = 4$, $l = 1$?
   (A) 2
   (B) 3
   (C) 4
   (D) 5
10. The electron configuration that is impossible is:
   (A) 1s$^2$ 2s$^2$ 2p$^6$
   (B) 1s$^2$ 2s$^2$ 2p$^3$
   (C) 1s$^2$ 2s$^2$ 2p$^6$ 3s$^1$
   (D) 1s$^2$ 2s$^2$ 2p$^6$ 3s$^2$
   (E) 1s$^2$ 2s$^2$ 2p$^6$ 2d$^2$

11. A neutral atom has an atomic number of 30 and a mass number of 62, the atom must contain:
   (A) 92 neutrons
   (B) 62 electrons
   (C) 29 neutrons
   (D) 30 electrons

12. Atom X has 12 protons, 12 electrons, and 13 neutrons. Atom Y has 10 protons, 10 electrons, and 15 neutrons. It can therefore be concluded that:
   (A) atoms X and Y are isotopes.
   (B) atom X is more massive than atom Y.
   (C) atoms X and Y have the same mass number.
   (D) atoms X and Y have the same atomic number.

13. Which set of quantum numbers (n, l, m$\_l$, m$\_s$) represents the outermost electron in a gaseous aluminum atom?
   (A) 2, 1, 0, +1/2
   (B) 2, 1, -1, +1/2
   (C) 3, 0, 0, +1/2
   (D) 3, 1, -1, +1/2

14. Which species is paramagnetic in the gaseous state?
   (A) Cu
   (B) Zn$^{2+}$
   (C) Sn$^{2+}$
   (D) Cr$^{3+}$

15. A neutral atom which has 42 electrons and a mass number of 93 has
   (A) an atomic number of 51.
   (B) a nucleus containing 51 neutrons.
   (C) a nucleus containing 40 neutrons.
   (D) a nucleus containing 51 protons.

16. A sodium ion, Na$^+$, contains the same number of electrons as
   (A) a sodium atom, Na.
   (B) a magnesium atom, Mg.
(C) a potassium ion, K⁺.
(D) a neon atom, Ne.

Electronic Structure Worksheet 2
1. If two atomic species are isotopes, then
   (A) both atoms must have identical nuclei.
   (B) the nuclei of both atoms contain the same number of neutrons.
   (C) the nuclei of both atoms contain the same number of protons.
   (D) both atoms must have the same mass.

2. What is the maximum number of electrons that can have a principal quantum number of 3 within one atom?
   (A) 3
   (B) 8
   (C) 18
   (D) 32

3. Which atom is paramagnetic in the gaseous state?
   (A) K
   (B) Ca
   (C) Zn
   (D) Kr

4. The partial symbol for a particular ion is ²⁶M²⁺. The number of electrons contained in one of these ions is
   (A) 2
   (B) 10
   (C) 12
   (D) 24

5. How many unpaired electrons are found in the most stable electronic state (ground state) of a sulfur atom?
   (A) 0
   (B) 2
   (C) 4
   (D) 6

6. The electron configuration for Mn²⁺ is:
   (A) [Ar]4s² 3d⁶
   (B) [Ar]3d⁵
   (C) [Ar]4s¹ 3d⁵
   (D) [Ar]4s¹ 3d⁴
7. Which set contains three isoelectronic species?
(A) Zn, Cd, Hg
(B) Br⁺, Kr, Rb⁻
(C) P²⁻, Se²⁻, I⁻
(D) F⁻, Na⁺, Mg²⁺

8. Which of the following refers to the ground-state electron configuration of an atom?
(A) 1s¹ 2s¹
(B) [Kr]5p¹
(C) [Ne]3s¹ 3p²
(D) [Ar]4s² 3d⁶

9. The maximum number of electrons in an atom that can have quantum numbers n = 2, l = 1 is:
(A) 2
(B) 6
(C) 8
(D) 4

10. Which of the following pairs contains isoelectronic species?
(A) Na and Mg⁺
(B) P⁻ and Se
(C) N²⁻ and Ne
(D) O²⁻ and Na⁺

11. Which species is diamagnetic in the ground state?
(A) N
(B) Zn²⁺
(C) Cu²⁺
(D) O⁻

12. An atom of iron-56, Fe⁵⁶, contains
(A) 26 electrons, 26 protons, 56 neutrons
(B) 56 electrons, 26 protons, 26 neutrons
(C) 56 electrons, 56 protons, 26 neutrons
(D) 26 electrons, 26 protons, 30 neutrons

13. What is the maximum number of electrons that can occupy the 5f subshell?
(A) 10
(B) 14
(C) 7
(D) 2

14. Ca⁴⁰, K³⁹, and Ti⁴² all have the same
(A) number of electrons.
(B) atomic number.
(C) mass number.
(D) number of neutrons.

15. Element X, whose atoms have an outer-shell electron configuration $\text{ns}^2 \text{np}^3$, is most likely to react chemically to form ions which have a charge of
(A) +3
(B) + 1
(C) -3
(D) -2

Electronic Structure Worksheet 3
1. What is the atomic number of the first element in the periodic table to have a filled d orbital?
(A) 30
(B) 2
(C) 2
(D) 29

2. Which atom or ion is given with an excited outer electron configuration?
(A) Na: 3s$^1$
(B) B: 2p$^3$
(C) He: 1s$^2$
(D) H$^{-1}$: 1s$^2$

3. Which of the following represents the ground state electron configuration of the oxygen atom?

$$
\begin{array}{c|ccc}
1s & 2s & 2p \\
\hline
(A) & \uparrow \downarrow & \uparrow \uparrow & \uparrow \\
(B) & \uparrow \downarrow & \uparrow \downarrow & \uparrow \uparrow \uparrow \downarrow \\
(C) & \uparrow \downarrow & \uparrow \downarrow & \uparrow \downarrow \uparrow \\
(D) & \uparrow \downarrow & \uparrow \downarrow & \uparrow \uparrow \\
\end{array}
$$

4. What is the wavelength of the radiation emitted from a mercury arc sunlamp if the frequency of the radiation is about $1.2 \times 10^{15}$ sec$^{-1}$? (c = $3.0 \times 10^{10}$ cm/sec)
(A) $2.0 \times 10^{-5}$ cm
(B) $4.0 \times 10^{-4}$ cm
(C) $2.5 \times 10^{-6}$ cm
(D) $2.5 \times 10^{-5}$ cm

5. How many unpaired electrons are there in the Ti$^{3+}$ ion?
(A) 0
(B) 2
6. The quantum numbers 3, 1, -1, +1/2
(A) refer to an electron in the p orbital of the 3rd shell.
(B) refer to an electron in the p orbital of the 2nd shell.
(C) refer to an electron in carbon.
(D) refer to an electron in the s orbital of the 3rd shell.

7. The charge on the nucleus of a Mg$^{2+}$ ion is
(A) +2
(B) +10
(C) +12
(D) –2

8. The sublevel that can be occupied by a maximum of 10 electrons is identified by the letter
(A) d
(B) f
(C) p
(D) s

9. An orbital may never be occupied by
(A) 1 electron
(B) 2 electrons
(C) 3 electrons
(D) 0 electrons

10. Which particle consists of 13 protons, 14 neutrons, and 10 electrons?
(A) neon atom
(B) sodium atom
(C) aluminum ion
(D) silicon atom

11. The number of orbitals in the 2nd shell of an atom is
(A) 1
(B) 9
(C) 16
(D) 4

12. The first element in the periodic table having the first completed p orbital is
(A) He
(B) Be
(C) O
(D) Ne
13. One of the outermost electrons in a strontium atom in the ground state can be described by which of the following sets of four quantum numbers?
(A) 5, 2, 0, ½
(B) 5, 1, 1, ½
(C) 5, 1, 0, ½
(D) 5, 0, 1, ½
(E) 5, 0, 0, ½

14. All of the following attempts to write a ground state electronic configuration are incorrect except for one. The correct one is
(A) [Ne]3s$^2$ 3p$^6$ 3d$^{10}$
(B) [He]2s$^2$ 2p$^3$
(C) [Ne]3p$^3$
(D) [Ne]3s$^2$ 3p$^4$

15. The energy levels of the hydrogen atom are related to the principle quantum number by:
(A) $E = k/n$
(B) $E = k/n^2$
(C) $E = kn$
(D) $E = k/n^2$

16. How many atomic orbital configurations that satisfy Hunds’ Rule can be written for the 1s$^2$ 2s$^2$ 2p$^2$ structure of the carbon atom such that all the p electrons have a spin quantum number of $+1/2$?
(A) 1
(B) 2
(C) 3
(D) 4

Electronic Structure Worksheet 4
1. The first line in the hydrogen spectrum is very difficult to see and it's the kind of light that causes skin cancer. It is a dark purple color with a wavelength of 410.18 nm. Calculate:

(A) the frequency of this light.

(B) the change in energy that produces this color.

\[ h = 6.626 \times 10^{-34} \text{ J} \cdot \text{s/particle} \]
\[ c = 2.998 \times 10^8 \text{ m/s} \]
2. Which of the following would have the higher ionization energy?
(A) K⁺ or Ca
(B) S²⁻ or Cl⁻
(C) O or S
(D) O or N
(E) P or O

3. Which of the following has the more metallic character?
(A) Pb or Rb
(B) At or Ar
(C) Na or Fr

4. Match the species in the left column with the isoelectronic ion or atom from the right.

(A) H⁻ ________
(B) Se²⁻ ________
(C) In⁺ ________
(D) Ar ________
(E) Zn²⁺ ________
(F) Na⁺ ________
(G) Cs⁺ ________

I. Ga³⁺
II. Al³⁺
III. Ti³⁺
IV. Ag⁺
V. P³⁻
VI. Ni
VII. Cd
VIII. Y³⁺
IX. Cu
X. Te⁵⁻
XI. Mn⁴⁺
XII. Be²⁺

5. Which atom has a ground state p³ electronic configuration?
(A) Mg
(B) Ga
(C) Al
(D) P

6. The ion that has the ground state electronic configuration of [Ne]3s² 3p³ is
(A) Si⁺¹
(B) Al⁺³
(C) S⁺¹
(D) Cl⁻¹

7. The number of orbitals in a d sublevel is
(A) 1
(B) 3
(C) 7
(D) 5

8. What is the correct set of quantum numbers for the highest energy electron of Na?
(A) 3, 0, 0, 1/2
(B) 3, 1, 0, 1/2
(C) 3, 1, -1, -1/2
(D) 1,0, 0, -1/2
9. An atom of element X absorbs a photon and an electron is transferred from the ground state to a higher energy level. Which of the following represents such a transition?
(A) n = 2 to n = 4.
(B) n = 3 to n = 1.
(C) n = 1 to n = -5
(D) n = 1 to n = 3.
(E) n = 0 to n = 2.

10. Which of the following sets of quantum numbers is not possible?
(A) 3, 1, 0, 1/2
(B) 1, 1, 0, -1/2
(C) 2, 1, 1, 1/2
(D) 4, 2, -2, -1/2
(E) 2, 1, 0, 1/2

11. Ni$^{2+}$ is isoelectronic with which of the following species?
(A) Mn
(B) Fe$^{1+}$
(C) V$^{3+}$
(D) Co$^{1+}$
(E) Ar

12. The exceptions to the Aufbau principle are explained by
(A) no two electrons can have the same quantum numbers.
(B) the merging of orbitals between two sublevels
(C) an increase in the stability of the electron configuration.
(D) the Bohr model's inability to deal with many-electron systems.

Electronic Structure Worksheet 5

1. Explain briefly why the emission spectrum of a hydrogen atom is comprised of discrete lines instead of a continuous broad band of emitted light.

2. List all the possible values of the angular quantum number, l, for a principle quantum number 4. Describe the difference between orbitals for each of the values of l.

3. There are simple mathematical relationships for the number of orbitals in a particular shell and the number of electrons in a shell. Derive them.
4. Elements may be synthetically created and it is theoretically probable that element 120, when created or perhaps discovered, will be more stable than the recent additions to the periodic table, elements 103 - 109. Basing your answer on the electronic configuration of the molecule, in which group will element 120 reside?

5. Identify the quantum number that specifies each of the following things.
(A) The spatial orientation of the orbital.
(B) The spin of the electrons that occupy the orbital.
(C) The size of the orbital.
(D) The shape of the orbital.

6. Atoms and molecules can emit and absorb light (electromagnetic radiation). Describe both of these processes and how they affect the electronic stability of atoms or molecules.

7. Give the symbols for the element of lowest atomic number whose ground state has:
(A) a completed p sublevel.
(B) four 4d electrons.
(C) five f electrons.
(D) two 3s electrons.
(E) eight electrons in its valence shell.

8. In the photoelectric effect, photons strike the surface of a metal and electrons are emitted from the metal. Describe why the photoelectric effect reinforces quantum theory and calculate the speed of an electron emitted from lithium by a photon with wavelength 420 nm.
\[ h = 6.63 \times 10^{-34} \text{ J•s} \quad \text{work function lithium} = 2.3 \text{ eV} \]

9. The hydrogen emission spectrum is characterized by several series of sharp emission lines in the infrared (Paschen series, Brackett series, etc.), visible (Balmer series) and ultraviolet (Lyman series) portions of the electromagnetic spectrum.
(A) How is the discrete nature of the emission spectrum of hydrogen explained in terms of the electronic energy levels of the hydrogen atom.
(B) Account for the existence of several series of lines in the spectrum. What quantity distinguishes one series of lines from another?

(C) Draw an electronic energy level diagram for the hydrogen atom and indicate on it the transition corresponding to the line of lowest frequency in the Lyman series.

(D) What is the difference between an absorption spectrum and an emission spectrum?

(E) At room temperature the absorption spectrum of the hydrogen atom exhibits only the transitions of the Lyman series. Explain this.

10. The Heisenberg Uncertainty Principle and deBroglie's ideas of the wave-particle duality of matter are two founding concepts for the quantum mechanical picture of electrons in atoms.
(A) State the Heisenberg uncertainty principle with respect to its determining the position and momentum of an object.

(B) What part of the Bohr theory of the atom is considered unrealistic as a result of the Heisenberg uncertainty principle?

(C) Explain why the Heisenberg uncertainty principle or the wave nature of particles is not a practical way of examining the behavior of macroscopic objects, but becomes most significant when describing the behavior of electrons or systems on a very small scale.
Electronic Structure Worksheet Answer Key

**WORKSHEET 1**

**WORKSHEET 2**

**WORKSHEET 3**

**WORKSHEET 4**
1a) $7.31 \times 10^{14} \text{ s}^{-1}$, b) $4.843 \times 10^{-19} \text{ J}$

2a) K$^+$, b) Cl$^-$, c) O, d) N, e) O

3a) Rb, b) At, c) Fr

4a) XII, b) VIII, c) VII, d) V, e) I, f) II, g) X

**WORKSHEET 5**
1) Energy is quantized: electrons can only have certain energies. When an electron makes a transition from a higher energy level to a lower energy level, the excess energy may be released in the form of light. The frequency of the light depends on the energy difference between the levels. Since electrons occupy only specific energy levels, only specific differences (and thus only certain frequencies) will result. So you see line spectra corresponding to those frequencies.

2) $n = 4$ allows for 4 values of $L$

3) $n^2$, $2n^2$

4) IIA (Alkaline earth metals)

5a) m l , b) m s , c) n, d) l

6) energy transitions, ionization energy, excited state stability

7a) Ne, b) Mo, c) Pm, d) Mg, e) Ne

8a) $5.8 \times 10^6 \text{ m/s}$ b) Photon energy is quantized. No electrons are ejected unless the photons have the threshold energy or greater.

9a) Same answer as question 1.

b) In a series, all transitions are from some higher energy level to the same final level. The final energy level distinguishes one series from another.

c) The lowest energy transition is equal to the lowest frequency transition. For the Lyman series, the final state is $n = 1$.

d) In an absorption spectrum, energy is absorbed by the system to raise electrons from lower levels to higher levels. The spectrum that results will have some part for the input light subtracted out – those frequencies corresponding to the energy level gaps of the system. An emission spectrum is energy being released in the form of light when electrons make transitions from higher levels to lower ones.

e) At room temperature, essentially all electrons are in the ground state. Thus they can make transitions from the ground state to higher states (Lyman series). There are no electrons in excited states so there are no transitions from excited state to excited state.
10) a) $\Delta x \Delta p \geq \hbar$  
   b) Electrons cannot be confined to specific orbits.  
   c) The mass(momentum) of macroscopic particles is so many orders of magnitude larger than that of Planks constant that the uncertainty principle is not relevant (although its still valid).