

## Electron Configuration Worksheet (and Lots More!!)

### Brief Instructions

An electron configuration is a method of indicating the arrangement of electrons about a nucleus. A typical electron configuration consists of numbers, letters, and superscripts with the following format:

1. A number indicates the energy level (The number is called the principal quantum number.).
2. A letter indicates the type of orbital; s, p, d, f.
3. A superscript indicates the number of electrons in the orbital. Example:  $1s^2$  means that there are two electrons in the 's' orbital of the first energy level. The element is helium.

To write an electron configuration:

1. Determine the total number of electrons to be represented.
2. Use the Aufbau process to fill the orbitals with electrons. The Aufbau process requires that electrons fill the lowest energy orbitals first. In another words, atoms are built from the ground upwards.
3. The sum of the superscripts should equal the total number of electrons.

Example:  ${}_{12}\text{Mg } 1s^2 2s^2 2p^6 3s^2$

### Configuration Writing Practice

Write a ground state electron configuration for each neutral atom. Ground state means that all of the lowest possible energy levels (up to the proper number of electrons for the element) are filled.

1. Na    2. Pb    3. Sr    4. U    5. N    6. Ag    7. Ti    8. Ce    9. Cl    10. Hg

Write a ground state electron configuration for these ions. Remember that ions have a change in the total number of electrons (positive have lost electrons and negative have gained).

Example:  $\text{N}^{3-}$  is  $1s^2 2s^2 2p^6$ . It has three extra electrons.

11.  $\text{O}^{2-}$     12.  $\text{Fe}^{2+}$     13.  $\text{B}^{3+}$     14.  $\text{Ni}^{2+}$     15.  $\text{K}^+$     16.  $\text{Co}^{3+}$

An excited atom has an electron or electrons which are not in the lowest energy state. Excited atoms are unstable energetically. The electrons eventually fall to a lower level. \* is used to indicate an excited atom. For example:  ${}^*\text{Li } 1s^2 3p^1$ . (The ground state for Li is  $1s^2 2s^1$ .)

Write an excited state electron configuration for each.    17. Al    18. Ar    19. K    20. C

21. If each orbital can hold a maximum of two electrons, how many electrons can each of the following hold?    a. 2s    b. 5p    c. 4f    d. 3d    e. 4d
22. What is the shape of an s orbital?
23. How many s orbitals can there be in an energy level?
24. How many electrons can occupy an s orbital?
25. What is the shape of a p orbital?
26. How many p orbitals can there be in an energy level?
27. Which is the lowest energy level that can have a s orbital?
28. Which is the lowest energy level that can have a p orbital?
29. Is it possible for two electrons in the same atom to have exactly the same set of quantum numbers?
30. Distinguish between an atom in its ground state and an excited atom.
31. How many d orbitals can there be in an energy level?
32. How many d electrons can there be in an energy level?
33. Which is the lowest energy level having d orbitals?
34. How many f electrons can there be in an energy level?
35. Which is the lowest energy level having f orbitals?

36. How many f orbitals can there be in an energy level?
37. How many energy levels are partially or fully occupied in a neutral atom of calcium?
38. Why do the fourth and fifth series of elements contain 18 elements, rather than 8 as do the second and third series?
39. Which sublevels of the 3rd energy level are filled (a) in the element argon (b) in the element krypton?
40. Why does it take more energy to remove an electron from  $\text{Al}^+$  than from  $\text{Al}$ ?
41. What does the term principal quantum number refer to?
42. What is meant by the electron configuration of an atom?
43. What is the maximum number of electrons that can be present in an atom having three principal energy levels?
44. Which of the following notations shows the electron configuration of a neutral atom in an excited state? Name the element, and explain how you know it is excited:
- (a)  $1s^2 2s^2 2p^1$       (b)  $1s^2 2s^2 2p^3 3s^1$       (c)  $1s^2 2s^2 2p^6 3s^2 3p^1$
45. Isoelectronic species have similar electron configurations. Which of these are isoelectronic?
- (a)  $\text{Li}^+$ ,  $\text{H}^-$ ,  $\text{He}$       (b)  $\text{Ca}^{2+}$ ,  $\text{Ne}$ ,  $\text{S}^{2-}$

46. For the following elements list the electron configuration. If there is no charge listed, assume it is neutral.

- |                            |                           |                           |                           |
|----------------------------|---------------------------|---------------------------|---------------------------|
| a. oxygen                  | b. cesium                 | c. krypton                | d. titanium               |
| e. scandium                | f. nitrogen               | g. chlorine               | h. fluorine <sup>1-</sup> |
| i. arsenic                 | j. francium               | k. selenium <sup>3-</sup> | l. copper <sup>1+</sup>   |
| m. potassium <sup>2-</sup> | n. antimony <sup>2+</sup> | o. thorium <sup>1-</sup>  | p. mercury <sup>2+</sup>  |

47. For the following elements list the shorthand electron configuration

- |                           |                          |               |                             |
|---------------------------|--------------------------|---------------|-----------------------------|
| a. boron                  | b. cadmium               | c. phosphorus | d. neon                     |
| e. radon                  | f. iodine                | g. strontium  | h. chromium <sup>3+</sup>   |
| j. nickel                 | k. iron                  | l. astatine   | m. molybdenum <sup>2-</sup> |
| n. rubidium <sup>3-</sup> | o. bromine <sup>1+</sup> | p. xenon      | q. europium <sup>3+</sup>   |

48. For the following electron configurations choose 3 possible elements (or ions) they may represent

- |    |  |
|----|--|
| a. | $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$   |
| b. | $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$   |
| c. | $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$   |
| d. | $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^1$              |
| e. | $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^{14} 6d^8$ |
| f. | $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{10}$                                |
| g. | $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^4$                   |
| h. | $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$   |
| i. | $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^2$   |
| j. | $[\text{Kr}] 5s^2 4d^{10} 5p^3$  |
| k. | $[\text{Kr}] 5s^2 4d^{10} 5p^6$  |
| l. | $[\text{Ar}] 4s^1$   |
| m. | $[\text{Xe}] 6s^2 4f^{10}$   |
| n. | $[\text{Xe}] 6s^2 4f^{14} 5d^7$  |
| o. | $[\text{Ne}] 3s^2 3p^1$  |

## Writing Electron Configurations

1. Electrons occupy the lowest energy orbital first, then move to the next one and so on. (The "Aufbau" Principle)
2. Orbitals are considered to be in the same shell if they have the same first number (no matter in what order filling is done).
3. An atom will gain or lose electrons in order to have eight electrons in its outer shell. (The "Octet" Rule)
4. The outer shell is the highest numbered shell which has electrons in it. Only s and p orbitals are part of the outer shell.

An atom has the tendency to lose electrons (to another atom) or to gain electrons (from another atom) in order to make the outer shell complete with eight electrons. Atoms with a complete outer shell (eight electrons) are considered stable. Some atoms naturally have eight electrons in their outer shell and are very stable. (Helium is the exception being stable with two electrons in its outer shell.) Complete the following chart:

Element	Atomic number	Number electrons in each E level	Electron Configuration	Number e <sup>-</sup> probably lost or gained	After loss or gain, # e <sup>-</sup> left	Charge on ion
O	8	2, 6	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>4</sup>	gain 2	10	-2
Na	11	2, 8, 1	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>6</sup> 3s <sup>1</sup>	lose 1	10	+1
S						
K						
Al						
Cl						
Xe						
Ca						
F						
Br						
N						
Ar						
I						
Sr						

## Worksheet Questions

1. Define wave crest, trough, amplitude, and wavelength.
2. Draw a wave and label it, using the following terms: crest, trough, amplitude, wavelength.
3. Draw a wave with a wavelength of 3.0 cm.
4. A cork floating on water moves up and down 10 times in 30 seconds. What is the frequency of the water wave?
5. A tuning fork produces a sound of musical note middle C. It moves back and forth 256 times each second. What is the frequency of the tuning fork?
6. What is the wavelength of sound waves having a frequency of  $256 \text{ sec}^{-1}$ . at  $20^\circ\text{C}$ ? Speed of sound = 340 m/sec
7. What is the frequency of a sound wave with a wavelength of 1 meter at  $20^\circ\text{C}$ ?
8. What is the frequency of sound waves which have wavelengths of 6 cm when the air temperature is  $20^\circ\text{C}$ ?
9. What is the frequency of sound waves which have wavelengths of 12 meters when the air temperature is  $20^\circ\text{C}$ ?
10. What is the electromagnetic spectrum?
11.
  - a. How are infrared waves different from red light waves?
  - b. How are ultraviolet waves different from violet light waves?
  - c. In what ways are infrared, red, ultra-violet, and violet light waves alike?
12. What is the wavelength in meters of a radio wave with a frequency of  $540,000 \text{ sec}^{-1}$ .
13. The electromagnetic waves used in FM broadcasting by radio or television have frequencies of approximately 100 megahertz. In standard AM radio broadcasting, the frequency is about 1 megahertz. (Remember mega =  $10^6$ .) Using the relationship  $\lambda\nu = c$ , calculate the wavelengths used in AM and FM broadcasting.
14. Choose your favorite radio station. Using its frequency ( $\nu$ ), calculate its wavelength ( $\lambda$ ). Be sure to give the call letters of the station.
15. For each of the following wavelengths of visible light, determine the frequency and identify the region of the electromagnetic spectrum to which it belongs.
  - a.  $2.0 \times 10^{-14}$  meter
  - b.  $4.0 \times 10^{-9}$
  - c.  $6.0 \times 10^{-7}$  meter
  - d. 1.00 meter
16. For each of the following frequencies of visible light, determine the wavelength, and identify the color of light associated with each frequency.
  - a.  $6.4 \times 10^{14} \text{ sec}^{-1}$
  - b.  $5.5 \times 10^{14} \text{ sec}^{-1}$
  - c.  $5.0 \times 10^{14} \text{ sec}^{-1}$

17. A chemist is using radiation with a frequency of  $6 \times 10^{13} \text{ sec}^{-1}$ .

a. What is the wavelength of this radiation in meters?

b. Identify this radiation as red, blue, infrared, ultraviolet, and so on.

c. Estimate the energy in kJ for one photon of this radiation.

Plank's constant ( $h$ ) is  $6.63 \times 10^{-34} \text{ J sec}$ . The velocity of light is  $3.00 \times 10^{10} \text{ meters/sec}$ . Use the relationship  $E = h\nu$ .

18. Use Plank's constant ( $6.63 \times 10^{-34} \text{ J sec}$ ) to determine the energy which corresponds to the following frequencies of the values by  $6.022 \times 10^{23}$  photons/mole.

a.  $4.567 \times 10^{14}$  hertz

b.  $6.165 \times 10^{14}$  hertz

c.  $6.905 \times 10^{14}$  hertz

d.  $7.307 \times 10^{14}$  hertz

e.  $7.550 \times 10^{14}$  hertz

19. In what region of the electromagnetic spectrum would you look for the spectral line resulting from the electronic transition from the fifth to the tenth electronic level in the hydrogen atom? ( $R_H = 1.10 \times 10^{-5} \text{ cm}^{-1}$ )

20. Use the Rydberg equation to calculate the frequency of a line in the hydrogen spectrum corresponding to a transition from  $n = 5$  to  $n = 4$ . Identify the spectral region or color which corresponds to this frequency.

