9.E: Electrons in Atoms and the Periodic Table (Exercises)

9.1: Blimps, Balloons, and Models of the Atom

9.2: Light: Electromagnetic Radiation

9.3: The Electromagnetic Spectrum

1. Choose the correct word for the following statement. Blue light has a (longer or shorter) wavelength than red light.
2. Choose the correct word for the following statement. Yellow light has a (higher or lower) frequency than blue light.
3. Choose the correct word for the following statement. Green light has a (larger or smaller) energy than red light.
4. If "light A" has a longer wavelength than "light B", then "light A" has _______________ "light B".

    (a) a lower frequency than

    (b) a higher frequency than

    (c) the same frequency as

5. If "light C" has a shorter wavelength than "light D", then "light C" has _______________ "light D".

    (a) a larger energy than

    (b) a smaller energy than

    (c) the same energy as

6. If "light E" has a higher frequency than "light F", then "light E" has __________________ "light F".

    (a) a longer wavelength than

    (b) a shorter wavelength than

    (c) the same wavelength as

7. If "light G" has a higher frequency than "light H", then "light G" has __________________ "light H".

    (a) a larger energy than

    (b) a smaller energy than

    (c) the same energy as

8. If "light J" has larger energy than "light K", then "light J" has __________________ "light K".

    (a) a shorter wavelength than

    (b) a longer wavelength than

    (c) the same wavelength as

9. Which of the following statements is true?

    (a) The frequency of green light is higher than the frequency of blue light and the wavelength of green light is longer than the wavelength of blue light.

    (b) The frequency of green light is higher than the frequency of blue light and the wavelength of green light is shorter than the wavelength of blue light.

    (c) The frequency of green light is lower than the frequency of blue light and the wavelength of green light is shorter than the wavelength of blue light.

    (d) The frequency of green light is lower than the frequency of blue light and the wavelength of green light is longer than the wavelength of blue light.

    (e) The frequency of green light is the same as the frequency of blue light and the wavelength of green light is shorter than the wavelength of blue light.

10. As the wavelength of electromagnetic radiation increases:

    (a) its energy increases.

    (b) its frequency increases.

    (c) its speed increases.

    (d) more than one of the above statements is true.

    (e) none of the above statements is true.

11. List three examples of electromagnetic waves.
12. Why do white objects appear white?
13. Name the colors present in white light in order of increasing frequency.
14. Why do objects appear black?

9.4: The Bohr Model: Atoms with Orbits

1. Decide whether each of the following statements is true or false:

    (a) Niels Bohr suggested that the electrons in an atom were restricted to specific orbits and thus could only have certain energies.

    (b) Bohr's model of the atom can be used to accurately predict the emission spectrum of hydrogen.

    (c) Bohr's model of the atom can be used to accurately predict the emission spectrum of neon.

    (d) According to the Bohr model, electrons have more or less energy depending on how far around an orbit they have traveled.

2. According to the Bohr model, electrons in an atom can only have certain, allowable energies. As a result, we say that the energies of these electrons are _______.
3. The Bohr model accurately predicts the emission spectra of atoms with…

    (a) less than 1 electron.

    (b) less than 2 electrons.

    (c) less than 3 electrons.

    (d) less than 4 electrons.

4. Consider an He^+ atom. Like the hydrogen atom, the He+ atom only contains 1 electron, and thus can be described by the Bohr model. Fill in the blanks in the following statements.

    (a) An electron falling from the n = 2 orbit of He^+ to the n = 1 orbit of He^+ releases ______ energy than an electron falling from the n = 3 orbit of He^+ to the n = 1 orbit of He^+.

    (b) An electron falling from the n = 2 orbit of He^+ to the n = 1 orbit of He^+ produces light with a ______ wavelength than the light produced by an electron falling from the n = 3 orbit of He^+ to the n = 1 orbit of He^+.

    (c) An electron falling from the n = 2 orbit of He^+ to the n = 1 orbit of He^+ produces light with a ______ frequency than the light produced by an electron falling from the n = 3 orbit of He^+ to the n = 1 orbit of He^+.

5. According to the Bohr model, higher energy orbits are located (closer to/further from) the atomic nucleus. This makes sense since negative electrons are (attracted to/repelled from) the positive protons in the nucleus, meaning it must take energy to move the electrons (away from/towards) the nucleus of the atom.
6. According to the Bohr model, what is the energy of an electron in the first Bohr orbit of hydrogen?
7. According to the Bohr model, what is the energy of an electron in the tenth Bohr orbit of hydrogen?
8. According to the Bohr model, what is the energy of an electron in the seventh Bohr orbit of hydrogen?
9. If an electron in a hydrogen atom has an energy of -6.06×10^{-20} J, which Bohr orbit is it in?
10. If an electron in a hydrogen atom has an energy of -2.69×10^{-20} J, which Bohr orbit is it in?
11. If an electron falls from the 5th Bohr orbital of hydrogen to the 3rd Bohr orbital of hydrogen, how much energy is released (you can give the energy as a positive number)?

12. If an electron falls from the 6th Bohr orbital of hydrogen to the 3rd Bohr orbital of hydrogen, what wavelength of light is emitted? Is this in the visible light range?

9.5: The Quantum-Mechanical Model: Atoms with Orbitals

9.6: Quantum-Mechanical Orbitals and Electron Configurations

1. Match each quantum number with the property that they describe.

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(a) n	i. shape
(b) ℓ	ii. orientation in space
(c) m_l	iii. number of nodes

2. A point in an electron wave where there is zero electron density is called a __________.
3. Choose the correct word in each of the following statements.

   (a) The (higher/lower) the value of n, the more nodes there are in the electron standing wave.

   (b) The (higher/lower) the value of n, the less energy the electron has.

   (c) The (more/less) energy the electron has, the more nodes there are in its electron standing wave.

4. Fill in the blank. For lower values of n, the electron density is typically found ________ the nucleus of the atom, while for higher values of n, the electron density is typically found __________the nucleus of the atom.
5. Circle all of the statements that make sense: Schrödinger discovered that certain quantities in the electron wave equation had to be integers, because when they weren't, the wave equation described waves which…

   (a) were discontinuous

   (b) were too small

   (c) were too long and narrow

   (d) were too short and fat

   (e) "doubled back" on themselves

6. What are the allowed values of ℓ for an electron standing wave with $n = 4$?
7. How many values of ℓ are possible for an electron standing wave with $n = 9$?
8. What are the allowed values of ml for an electron standing wave with $\ell = 3$?
9. How many different orientations are possible for an electron standing wave with $\ell = 4$?
10. What are the allowed values of m_l for $n = 2$?
11. Fill in the blanks. When $\ell = 0$, the electron orbital is _________ and when $\ell = 1$, the electron orbital is _________ shaped.
12. The $n = 1$ s orbital has _____ nodes.
13. The $n = 2$ s orbital has _____ nodes.
14. The $n = 2$ p orbital has _____ nodes.
15. The $n = 1$ p orbital has _____ nodes.
16. There are ____ different p orbitals.
17. What energy level (or value of n) has s, p and d orbitals, but no f orbitals?
18. How many different d orbital orientations are there?
19. How many f orbital orientations are there?
20. How many different orbitals are there in the $n = 3$ energy level?

1. Write the electron configuration for beryllium. Beryllium has 4 electrons.
2. Write the electron configuration for silicon. Silicon has 14 electrons.
3. Write the electron configuration for nitrogen. Nitrogen has 7 electrons.
4. Write the electron configuration for chromium. Chromium has 24 electrons.
5. Write the electron configuration for silver. Silver has 47 electrons.

9.7: Electron Configurations and the Periodic Table

1. Use the Periodic Table to determine the energy level of the valence electrons in each of the following elements.

   (a) B
   (b) Ga
   (c) Rb
   (d) At
   (e) He

2. Fill in the blanks:

   (a) B is in the __ level block of the Periodic Table
   (b) Sr is in the __ level block of the Periodic Table
   (c) Fe is in the __ level block of the Periodic Table
   (d) Cs is in the __ level block of the Periodic Table
   (e) O is in the __ level block of the Periodic Table

3. Use the Periodic Table to determine the energy level and sublevel of the highest energy electrons in each of the following elements:

   (a) N
   (b) Ca
   (c) Rb
   (d) P
   (e) In

4. Decide whether each of the following statements is true or false.

   (a) Li has valence electrons in the n = 1 energy level.
   (b) Si has valence electrons in the n = 3 energy level.
   (c) Ga has valence electrons in the n = 3 energy level.
   (d) Xe has valence electrons in the n = 5 energy level.
   (e) P has valence electrons in the n = 2 energy level.

5. Match the element to the sublevel block it is found in:

Match the element to the sublevel block it is found in:

(a) C	i. *s* sublevel block
(b) Cs	ii. *p* sublevel block
(c) Ce	iii. *d* sublevel block
(d) Cr	iv. *f* sublevel block

6. The first row of the Periodic Table has:

   (a) 1 element
   (b) 2 elements
   (c) 3 elements
   (d) 4 elements
   (e) 5 elements

7. Use the Periodic Table to determine which of the following elements has the highest energy valence electrons.

(a) Sr
(b) As
(c) H
(d) At
(e) Na

8. Use the Periodic Table to determine which of the following elements has the lowest energy valence electrons.

(a) Ga
(b) B
(c) Cs
(d) Bi
(e) Cl

9. Which energy level does the first row in the *d* sublevel block correspond to?

9.8: The Explanatory Power of the Quantum-Mechanical Model

9.9: Periodic Trends: Atomic Size, Ionization Energy, and Metallic Character

1. Why is the atomic size considered to have "no definite boundary"?
2. How is atomic size measured?

(a) using a spectrophotomer
(b) using a tiny ruler (called a nano ruler)
(c) indirectly
(d) directly

3. Draw a visual representation of the atomic radii of an iodine molecule.
4. Which of the following would be smaller?

(a) In or Ga
(b) K or Cs
(c) Te or Po

5. Explain in your own words why Iodine is larger than Bromine.
6. What three factors determine the trend of atomic size going down a group?
7. What groups tend to show this trend?
8. Which of the following would have the largest atomic radii?

(a) Si
(b) C
(c) Sn
(d) Pb

9. Which of the following would have the smallest atomic radius?

(a) $2s^2$
(b) $4s^24p^3$
(c) $2s^22p^4$
(d) $4s^2$

10. Arrange the following in order of increasing atomic radii: Tl, B, Ga, Al, In.
11. Arrange the following in order of increasing atomic radii: Ge, Sn, C.
12. Which of the following would be larger?

    (a) Rb or Sn

    (b) Ca or As

13. Place the following in order of increasing atomic radii: Mg, Cl, S, Na.
14. Describe the atomic size trend for the rows in the Periodic Table.
15. Draw a visual representation of the periodic table describing the trend of atomic size.
16. Which of the following would have the largest atomic radii?

    (a) Sr

    (b) Sn

    (c) Rb

    (d) In

17. Which of the following would have the smallest atomic radii?

    (a) K

    (b) Kr

    (c) Ga

    (d) Ge

18. Place the following elements in order of increasing atomic radii: In, Ca, Mg, Sb, Xe.
19. Place the following elements in order of decreasing atomic radii: Al, Ge, Sr, Bi, Cs.
20. Knowing the trend for the rows, what would you predict to be the effect on the atomic radius if an atom were to gain an electron? Use an example in your explanation.
21. Knowing the trend for the rows, what would you predict to be the effect on the atomic radius if the atom were to lose an electron? Use an example in your explanation.

Ionization Energy

1. Define ionization energy and show an example ionization equation.
2. Draw a visual representation of the periodic table describing the trend of ionization energy.
3. Which of the following would have the largest ionization energy?

    (a) Na

    (b) Al

    (c) H

    (d) He

4. Which of the following would have the smallest ionization energy?

    (a) K

    (b) P

    (c) S

    (d) Ca

5. Place the following elements in order of increasing ionization energy: Na, O, Ca, Ne, K.
6. Place the following elements in order of decreasing ionization energy: N, Si, S, Mg, He.
7. Using experimental data, the first ionization energy for an element was found to be 600 kJ/mol. The second ionization energy for the ion formed was found to be 1,800 kJ/mol. The third ionization energy for the ion formed was found to be 2,700 kJ/mol. The fourth ionization energy for the ion formed was found to be 11,600 kJ/mol. And finally the fifth ionization energy was

found to be 15,000 kJ/mol. Write the reactions for the data represented in this question. To which group does this element belong? Explain.

8. Using electron configurations and your understanding of ionization energy, which would you predict would have higher second ionization energy: Na or Mg?
9. Comparing the first ionization energy (IE_1) of calcium, Ca, and magnesium, Mg, :

   (a) Ca has a higher IE_1 because its radius is smaller.
   (b) Mg has a higher IE_1 because its radius is smaller.
   (c) Ca has a higher IE_1 because its outer sub-shell is full.
   (d) Mg has a higher IE_1 because its outer sub-shell is full.
   (e) they have the same IE_1 because they have the same number of valence electrons.

10. Comparing the first ionization energy (IE_1) of beryllium, Be, and boron, B:

   (a) Be has a higher IE_1 because its radius is smaller.
   (b) B has a higher IE_1 because its radius is smaller.
   (c) Be has a higher IE_1 because its *s* sub-shell is full.
   (d) B has a higher IE_1 because its *s* sub-shell is full.
   (e) They have the same IE_1 because B has only one more electron than Be.

Electron Affinity

1. Define electron affinity and show an example equation.
2. Choose the element in each pair that has the lower electron affinity:

   (a) Li or N
   (b) Na or Cl
   (c) Ca or K
   (d) Mg or F

3. Why is the electron affinity for calcium much higher than that of potassium?
4. Draw a visual representation of the periodic table describing the trend of electron affinity.
5. Which of the following would have the largest electron affinity?

   (a) Se
   (b) F
   (c) Ne
   (d) Br

6. Which of the following would have the smallest electron affinity?

   (a) Na
   (b) Ne
   (c) Al
   (d) Rb

7. Place the following elements in order of increasing electron affinity: Te, Br, S, K, Ar.
8. Place the following elements in order of decreasing electron affinity: S, Sn, Pb, F, Cs.
9. Describe the trend that would occur for electron affinities for elements in Period 3. Are there any anomalies? Explain.
10. Comparing the electron affinity (EA) of sulfur, S, and phosphorus, P:

   (a) S has a higher EA because its radius is smaller.
   (b) P has a higher EA because its radius is smaller.

(c) S has a higher EA because its *p* sub-shell is half full.

(d) P has a higher EA because its *p* sub-shell is half full.

(e) they have the same EA because they are next to each other in the Periodic Table.
