

Chem 10 Chapter 9

Electrons in Atoms and the Periodic Table

Additional Practice Questions

Worked solutions at the end

Use your course periodic table and a calculator as needed. Show the setup, include units, and round at the end. Question numbers match the numbers in the solutions. Use heading navigation to move between sections.

In orbital diagrams, each bracketed group represents one orbital. Up and down arrows represent opposite spins. Use ground-state configurations unless stated otherwise.

Useful constants (standard rounded values):

$$c = 2.998 \times 10^8 \text{ m/s}; \quad h = 6.626 \times 10^{-34} \text{ J s}$$

Additional Practice Questions

Question 1

Green light can have a wavelength of 543 nm. What is the energy of one photon of this light?

- A) $1.08 \times 10^{-31} \text{ J}$
- B) $5.43 \times 10^{-7} \text{ J}$
- C) $3.66 \times 10^{-19} \text{ J}$
- D) $5.52 \times 10^{14} \text{ J}$
- E) $2.73 \times 10^{18} \text{ J}$

Question 2

Which observation most clearly supports the conclusion that electrons have wave properties?

- A) the emission spectrum of hydrogen
- B) the photoelectric effect
- C) scattering of alpha particles by metal foil
- D) diffraction
- E) cathode rays

Question 3

Which observation most clearly supports the conclusion that electrons in atoms have quantized energies?

- A) the emission spectrum of hydrogen
- B) the photoelectric effect
- C) scattering of alpha particles by metal foil
- D) diffraction
- E) cathode rays

Question 4

A photon with wavelength 500 nm supplies just enough energy to promote an electron from the second quantum level to the third in a particular element. What can be deduced?

- A) the energy of the $n = 2$ level
- B) the energy of the $n = 3$ level
- C) the sum of the energies of $n = 2$ and $n = 3$
- D) the difference in energy between $n = 2$ and $n = 3$
- E) all of these

Question 5

Use the hydrogen energy levels below. Which transition emits light with the longest wavelength?

$$E_4 = -0.1361 \times 10^{-18} \text{ J}$$

$$E_3 = -0.2420 \times 10^{-18} \text{ J}$$

$$E_2 = -0.5445 \times 10^{-18} \text{ J}$$

$$E_1 = -2.178 \times 10^{-18} \text{ J}$$

- A) $n = 4$ to $n = 3$
- B) $n = 4$ to $n = 2$
- C) $n = 4$ to $n = 1$
- D) $n = 3$ to $n = 2$
- E) $n = 2$ to $n = 1$

Question 6

An electron moves from 2p to 2s in a particular atom and emits a photon with wavelength 629.1 nm. What is the magnitude of the energy difference between the orbitals?

- A) 3.16×10^{-28} J
- B) 3.16×10^{-19} J
- C) 3.16×10^{-17} J
- D) 1.25×10^{-31} J
- E) none of these

Question 7

How many orbitals are possible in the $n = 2$ shell?

- A) 3
- B) 4
- C) 2
- D) 8
- E) 6

Question 8

Consider the representation of a 2p orbital shown below. Which statement best describes an electron in this orbital?

- A) The electron moves along the outer surface in a figure-eight path.
- B) The electron moves inside the two lobes but never outside their pictured surfaces.
- C) The electron is concentrated at the central node between the lobes.
- D) The electron moves in only one lobe at any given time.
- E) The electron's exact movement cannot be determined.

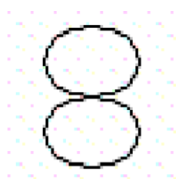


Figure description: Two rounded lobes arranged vertically, meeting at the center, represent a 2p orbital. The two lobes have the same orientation along a vertical axis.

Question 9

Which ground-state atom or ion has three unpaired electrons?

- A) N
- B) O
- C) Al
- D) S^{2-}
- E) Ti^{2+}

Question 10

Which is the complete ground-state electron configuration of tin?

- A) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5d^{10} 5p^2$
- B) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4d^{10} 4p^2$
- C) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 4p^6 5s^2 4d^{10} 5d^{10} 5p^2$
- D) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^2$
- E) none of these

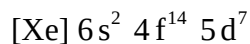
Question 11

The lowest-energy arrangement in a set of equal-energy orbitals has the maximum number of singly occupied orbitals permitted by the Pauli principle, with parallel spins. What is this rule called?

- A) the Aufbau principle
- B) Hund's rule
- C) the Heisenberg uncertainty principle
- D) the Pauli exclusion principle
- E) the quantum model

Question 12

An element has the configuration shown. Which class of elements does it belong to?



- A) transition elements
- B) alkaline earth elements
- C) halogens
- D) rare earth elements
- E) none of the above

Question 13

How many electrons are unpaired in the d orbitals of a ground-state iron atom?

- A) one
- B) two
- C) three
- D) four
- E) none of these

Question 14

Fluorine has nine electrons. How many are in s orbitals?

- A) 2
- B) 4
- C) 6
- D) 8
- E) none

Question 15

Which element has occupied d orbitals in its ground-state neutral atoms?

- A) Ba
- B) Ca
- C) Si
- D) P
- E) Cl

Question 16

Which element has a ground-state electron configuration that differs from the simplest Aufbau prediction?

- A) Ca
- B) Sc
- C) Ti
- D) V
- E) Cr

Question 17

Which process has an energy change equal to the first ionization energy of bromine?

- A) $\text{Br(s)} \rightarrow \text{Br}^+(\text{g}) + \text{e}^-$
- B) $\text{Br(l)} \rightarrow \text{Br}^+(\text{g}) + \text{e}^-$
- C) $\text{Br(g)} \rightarrow \text{Br}^+(\text{g}) + \text{e}^-$
- D) $\text{Br(s)} \rightarrow \text{Br}^+(\text{s}) + \text{e}^-$
- E) $\text{Br}_2(\text{g}) \rightarrow \text{Br}_2^+(\text{g}) + \text{e}^-$

Question 18

Order S, Cl, and F by increasing first ionization energy.

- A) S, Cl, F
- B) Cl, F, S
- C) F, S, Cl
- D) F, Cl, S
- E) S, F, Cl

Question 19

Order S, Cl, and F by increasing atomic radius.

- A) S, Cl, F
- B) Cl, F, S
- C) F, S, Cl
- D) F, Cl, S
- E) S, F, Cl

Question 20

Which atom has the largest first ionization energy?

- A) O
- B) Li
- C) Ne
- D) Be
- E) K

Question 21

Removing one electron from a neutral gas-phase sodium atom is what kind of process, and removing one from a neutral gas-phase fluorine atom is what kind of process?

- A) endothermic; exothermic
- B) exothermic; endothermic
- C) endothermic; endothermic
- D) exothermic; exothermic
- E) more information is needed

Question 22

Which comparison of krypton's 1s orbital with helium's 1s orbital is correct?

- A) Kr's 1s orbital is smaller because its nuclear charge draws electrons closer.
- B) Kr's 1s orbital is larger because it has more electrons.
- C) Kr's 1s orbital is smaller because its p and d orbitals crowd the s orbitals.
- D) The 1s orbitals have the same size because each can hold two electrons.
- E) Kr's 1s orbital is larger because its first ionization energy is lower.

Question 23

The two waves represent electromagnetic radiation in vacuum. Each spans 1.6×10^{-3} m, as indicated in the diagram.

- (a) Which has the longer wavelength? Calculate it.
- (b) Which has the higher frequency and larger photon energy? Calculate both for that wave.
- (c) Which has greater speed?
- (d) What type of electromagnetic radiation is each?

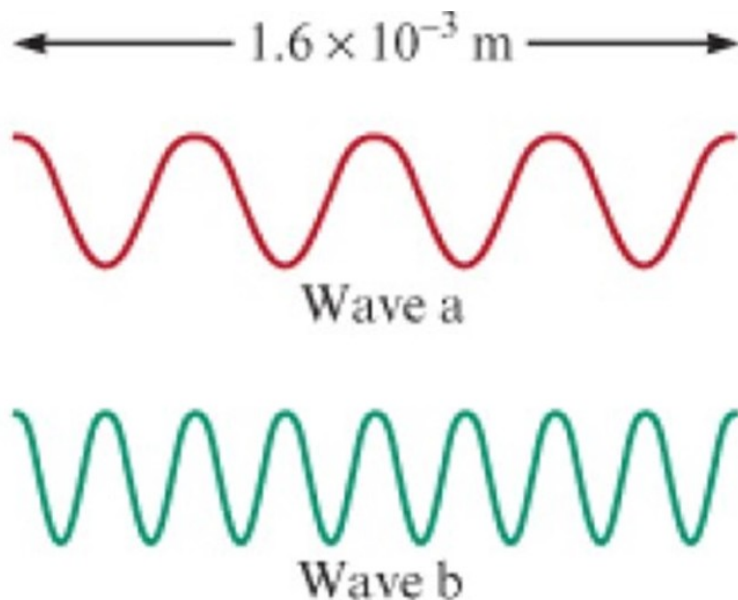


Figure description: Two waves extend across the same horizontal distance of 1.6×10^{-3} meters. The upper wave, labeled a, has four complete cycles. The lower wave, labeled b, has eight complete cycles. Wave a is drawn red and wave b green; the labels and cycle counts distinguish them without color.

Question 24

One radiation has frequency 107.1 MHz; another has wavelength 2.12×10^{-10} m; a third has photon energy 3.97×10^{-19} J. Identify each type and put them in order of increasing photon energy and increasing frequency.

Question 25

A carbon sample emits 1.98×10^5 J of light at wavelength 150. nm. Calculate the number of carbon atoms in the sample, assuming each atom emits one photon.

Question 26

Removing 1 mol of electrons from the surface of solid rubidium requires 208.4 kJ.

- (a) How much energy removes one electron?
- (b) What is the maximum wavelength of light capable of removing it?

Question 27

Use an electron mass of 9.11×10^{-31} kg. Calculate the de Broglie wavelength for:

- (a) an electron moving at 10.% of the speed of light
- (b) a 55 g tennis ball moving at 35 m/s

Question 28

A hydrogen electron is excited from the $n = 1$ ground state to $n = 3$. State whether each statement is true or false and correct the false ones.

- (a) It takes more energy to ionize the electron from $n = 3$ than from the ground state.
- (b) The electron is farther from the nucleus on average in $n = 3$ than in $n = 1$.
- (c) Light emitted in a 3-to-2 transition has shorter wavelength than light emitted in a 3-to-1 transition.
- (d) The wavelength emitted in a 3-to-1 transition equals the wavelength absorbed in a 1-to-3 transition.
- (e) For $n = 3$, the electron is in the first excited state.

Question 29

Identify each atom or ion from the outer-orbital diagram and description. Brackets represent individual orbitals; $\uparrow$ and $\downarrow$ denote opposite spins. Filled lower subshells are understood.

- (a) A neutral ground-state atom:

3s $[\uparrow\downarrow]$ 3p $[\uparrow\downarrow]$ $[\uparrow]$ $[\uparrow]$

- (b) A neutral excited-state atom, with two electrons in 1s:

2s $[\uparrow]$ 2p $[\uparrow\downarrow]$ $[\uparrow]$ $[\uparrow]$

- (c) A ground-state ion of charge -1 :

4s $[\uparrow\downarrow]$ 4p $[\uparrow\downarrow]$ $[\uparrow\downarrow]$ $[\uparrow]$

Question 30

How many unpaired electrons are in each ground-state species: O, O^+ , O^- , Os, Zr, S, F, and Ar?

Question 31

Arrange each group of atoms in order of increasing atomic size.

- (a) Te, S, Se
- (b) K, Br, Ni
- (c) Ba, Si, F

Question 32

Arrange each group of atoms in order of increasing first ionization energy.

(a) Te, S, Se

(b) K, Br, Ni

(c) Ba, Si, F

Solutions

Find the solution with the same question number.

Question 1 Solution

C. Convert the wavelength to meters, then calculate frequency and photon energy.

$$\lambda = 543 \text{ nm} \times \frac{10^{-9} \text{ m}}{1 \text{ nm}} = 5.43 \times 10^{-7} \text{ m}$$

$$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \text{ m/s}}{5.43 \times 10^{-7} \text{ m}} = 5.52 \times 10^{14} \text{ s}^{-1}$$

$$E(\text{photon}) = h\nu = (6.626 \times 10^{-34} \text{ J s})(5.52 \times 10^{14} \text{ s}^{-1})$$

$$E(\text{photon}) = 3.66 \times 10^{-19} \text{ J}$$

Question 2 Solution

D. Diffraction is fundamentally a wave property. Particles do not diffract when acting as classical particles; diffraction shows their wave behavior.

Question 3 Solution

A. The emission spectrum of hydrogen comes from light generated when an electron jumps from a higher level to a lower level. Since the spectrum is composed of discrete frequencies, rather than a continuous “rainbow” of frequencies like sunlight, the jumps can have only certain energy values. This in turn implies that the individual energy levels have only certain energy values.

Question 4 Solution

D. Photons emitted or absorbed by an atom correspond to a difference between energy levels. If the photon promotes the electron from the second quantum level to the third, its energy equals the difference between the $n = 2$ and $n = 3$ levels. It does not by itself give the separate energies of those levels.

$$E_3 - E_2 = \frac{hc}{\lambda}$$

Question 5 Solution

A. Longest wavelength means lowest frequency and lowest photon energy. For emission, photon energy is the positive magnitude of the atom's energy loss.

$$E(\text{photon}) = -\Delta E(\text{atom})$$

Directly from the supplied energy table, the 4-to-3 gap is $0.1059 \times 10^{-18} \text{ J}$, smaller than any of the other listed gaps.

An equivalent way to compare these hydrogen levels is the inverse-square method. For hydrogen, photon energy is proportional to the difference between the inverse squared principal numbers. The smallest difference gives the longest wavelength. In these choices, the adjacent levels with the largest n values are closest in energy: $n = 4$ and $n = 3$.

Or simply calculate the factor for every choice and find the smallest:

$$4 \rightarrow 3: \frac{1}{3^2} - \frac{1}{4^2} = 0.0486$$

$$4 \rightarrow 2: \frac{1}{2^2} - \frac{1}{4^2} = 0.1875$$

$$4 \rightarrow 1: \frac{1}{1^2} - \frac{1}{4^2} = 0.9375$$

$$3 \rightarrow 2: \frac{1}{2^2} - \frac{1}{3^2} = 0.1389$$

$$2 \rightarrow 1: \frac{1}{1^2} - \frac{1}{2^2} = 0.7500$$

The 4-to-3 transition has the smallest factor and hence the longest wavelength.

Question 6 Solution

B. Because the photon is emitted, the atom's energy change is negative. Photon energy is always positive.

$$\Delta E(\text{atom}) = -E(\text{photon}) = -\frac{hc}{\lambda}$$

$$\Delta E(\text{atom}) = -\frac{(6.626 \times 10^{-34} \text{ J s})(3.00 \times 10^8 \text{ m/s})}{629.1 \times 10^{-9} \text{ m}}$$

$$\Delta E(\text{atom}) \approx -3.16 \times 10^{-19} \text{ J}$$

The question asks for the energy difference, so we report its magnitude: $3.16 \times 10^{-19} \text{ J}$.

Question 7 Solution

B. For $n = 2$, we can have s and p orbitals. There is one 2s orbital and three 2p orbitals, giving $1 + 3 = 4$ orbitals.

Question 8 Solution

E. Orbitals have characteristic “shapes,” in the sense of where the electron is likely to spend time. However, they don’t have borders (any more than a fog has a border). Electrons are not confined to a surface or a curve.

In the case of a p orbital, they spend most of their time in the region containing most of the “probability density” and avoid the nodal plane between the two lobes. Since the electron is acting as a wave in the orbital, it is delocalized over the entire orbital, not just one “lobe” or the other. The “movement” of the electron cannot be exactly determined.

Question 9 Solution

A. For most of these neutral atoms and anions, start by examining the last subshell being filled. For a transition-metal cation, also account for the outer-shell electrons lost during ion formation.

N: $2p \uparrow \uparrow \uparrow$ — 3 unpaired

O: $2p \uparrow \downarrow \uparrow \uparrow$ — 2 unpaired

Al: $3p \uparrow \uparrow \uparrow$ — 1 unpaired

S^{2-} : $3p \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow$ — 0 unpaired

For Ti^{2+} , the two 4s electrons of Ti are lost on ionization, leaving $[Ar] 3d^2$.

Ti^{2+} : $3d \uparrow \uparrow \uparrow \uparrow \uparrow \uparrow$ — 2 unpaired

Nitrogen is the choice with three unpaired electrons.

Question 10 Solution

D. Tin has atomic number 50. The filled krypton core accounts for 36 electrons, followed by $5s^2$, $4d^{10}$, and $5p^2$.

Sn: $[Kr] 5s^2 4d^{10} 5p^2$

Sn: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^2$

The occupancies add to 50; no 5d electrons are needed.

Question 11 Solution

B. Hund’s rule. Fill equal-energy orbitals singly with parallel spins before pairing electrons. This differs from the Aufbau principle, which concerns filling lower-energy orbitals first.

Question 12 Solution

A. The 5d subshell is incomplete, so this is a transition element. The completed 4f subshell belongs to the core; the partially occupied 5d subshell determines the block classification.

Question 13 Solution

D. Iron has [Ar] 4s² 3d⁶. Its 3d subshell is populated as follows:

3d: $[\uparrow\downarrow] [\uparrow] [\uparrow] [\uparrow] [\uparrow]$

There are four unpaired electrons.

Question 14 Solution

B. Fluorine has two electrons in 1s and two in 2s, for a total of four electrons in s orbitals.

F: 1s² 2s² 2p⁵

Question 15 Solution

A. The “ground state” of an atom is the configuration obtained by applying the Aufbau principle as usual, filling the orbitals in the order of increasing energy as we move along the periodic table.

Ca, Si, P, and Cl are placed before we encounter the d block. We encounter the d block and start putting electrons (and keep them there) in the fourth period, after filling the 4s orbitals. Ca is the heaviest atom without any d electrons. Elements coming after that have d electrons, whether they happen to be in the d block or not. Ba has full 3d and 4d subshells.

Question 16 Solution

E. Cu and Cr are the two exceptions you are expected to remember. Chromium has [Ar] 4s¹ 3d⁵ rather than the simple prediction [Ar] 4s² 3d⁴. This gives a half-filled 3d subshell.

Question 17 Solution

C. First ionization energy refers to removing one electron from a neutral atom in the gas phase. The resulting ion and the removed electron are also gaseous species.

**Question 18 Solution**

A. Ionization energy generally increases from left to right across a period and decreases downward in a group. F is highest among S, Cl, and F because it is highest and farthest right. Cl is below F but to the right of S, so it is second. S is lowest.

Remember to rank them in increasing order:

**Question 19 Solution**

D. Use the radius trend, which runs opposite to the general first-ionization-energy trend. F is smallest because it is above Cl; Cl is smaller than S because it is to the right of S in the same period.



Question 20 Solution

C. Ionization energy generally increases as we go to the right and up in the periodic table. Li, Be, and K are far to the left of O and Ne.

We normally don't think about the noble gases, but Ne is to the right of O, and so it has a higher ionization energy. After all, noble gases must strongly resist ionization to deserve their name! If Ne formed a cation, it could just pair up with an anion and form a compound. That, it doesn't ordinarily do!

Question 21 Solution

C. A neutral gas-phase atom losing an electron requires energy, so ionization is endothermic. It is endothermic for both Na and F. This should not be confused with an atom gaining an electron.

Question 22 Solution

A. Krypton's 1s orbital is smaller than helium's 1s orbital because krypton's nuclear charge draws the electrons closer.

The 2s, 2p, 3s, 3p, 4s, 3d, and 4p electrons are almost entirely outside the 1s orbital (farther away from the nucleus) and offer little shielding of the extra protons that the Kr nucleus has (36 versus the two that He has). Thus the 1s orbital of Kr is attracted to a much larger effective nuclear charge and is pulled in a lot closer to the nucleus; therefore it has a smaller size than a He 1s orbital.

Question 23 Solution

Wavelength is the distance between consecutive wave peaks. Wave a shows four wavelengths and wave b shows eight.

$$\text{Wave a: } \lambda = \frac{1.6 \times 10^{-3} \text{ m}}{4} = 4.0 \times 10^{-4} \text{ m}$$

$$\text{Wave b: } \lambda = \frac{1.6 \times 10^{-3} \text{ m}}{8} = 2.0 \times 10^{-4} \text{ m}$$

Wave a has the longer wavelength. Frequency and photon energy are both inversely proportional to wavelength, so wave b has the higher frequency and larger photon energy.

$$\text{Wave b: } \nu = \frac{2.998 \times 10^8 \text{ m/s}}{2.0 \times 10^{-4} \text{ m}} = 1.5 \times 10^{12} \text{ s}^{-1}$$

$$\text{Wave b: } E = \frac{(6.626 \times 10^{-34} \text{ J s})(2.998 \times 10^8 \text{ m/s})}{2.0 \times 10^{-4} \text{ m}} = 9.9 \times 10^{-22} \text{ J}$$

Both waves travel at the same speed, c , the speed of light in vacuum. Both are infrared radiation.

Question 24 Solution

The 2.12×10^{-10} m radiation is X-rays. For 107.1 MHz radiation, calculate wavelength.

$$\lambda = \frac{c}{\nu} = \frac{2.9979 \times 10^8 \text{ m/s}}{107.1 \times 10^6 \text{ s}^{-1}} = 2.799 \text{ m}$$

This is FM radio radiation. For the stated photon energy, calculate wavelength from $E = hc/\lambda$.

$$\lambda = \frac{(6.626 \times 10^{-34} \text{ J s})(2.998 \times 10^8 \text{ m/s})}{3.97 \times 10^{-19} \text{ J}} = 5.00 \times 10^{-7} \text{ m}$$

This is visible green light. Photon energy and frequency have the reverse order to wavelength because both are inversely related to wavelength.

FM radio < visible green light < X-rays

This is the order of increasing frequency and increasing photon energy. Radio has the longest wavelength; X-rays have the shortest.

Question 25 Solution

First calculate the energy per photon.

$$E(\text{photon}) = \frac{hc}{\lambda}$$

$$E(\text{photon}) = \frac{(6.626 \times 10^{-34} \text{ J s})(2.998 \times 10^8 \text{ m/s})}{150. \text{ nm} \times \frac{1 \text{ m}}{10^9 \text{ nm}}}$$

$$E(\text{photon}) = 1.32 \times 10^{-18} \text{ J}$$

Use the total energy and the one-photon-per-atom assumption.

$$1.98 \times 10^5 \text{ J} \times \frac{1 \text{ photon}}{1.32 \times 10^{-18} \text{ J}} \times \frac{1 \text{ C atom}}{1 \text{ photon}} = 1.50 \times 10^{23} \text{ C atoms}$$

Question 26 Solution

Convert the energy per mole to energy per electron.

$$\frac{208.4 \text{ kJ}}{1 \text{ mol electrons}} \times \frac{1 \text{ mol electrons}}{6.0221 \times 10^{23} \text{ electrons}}$$

$$= 3.461 \times 10^{-22} \text{ kJ/electron} = 3.461 \times 10^{-19} \text{ J/electron}$$

At the maximum wavelength, a photon has just the required threshold energy. Carry the unrounded energy into this calculation.

$$E = \frac{hc}{\lambda}, \quad \lambda = \frac{hc}{E}$$

$$\lambda = \frac{(6.6261 \times 10^{-34} \text{ J s})(2.9979 \times 10^8 \text{ m/s})}{3.4605868 \times 10^{-19} \text{ J}}$$

$$\lambda = 5.740 \times 10^{-7} \text{ m} = 574.0 \text{ nm}$$

Longer wavelengths have less photon energy and cannot remove an electron by this one-photon process.

Question 27 Solution

(a) Find the speed, then use the electron mass in kilograms.

$$v = 0.10 \times 3.00 \times 10^8 \text{ m/s} = 3.0 \times 10^7 \text{ m/s}$$

$$\lambda = \frac{h}{mv} = \frac{6.63 \times 10^{-34} \text{ J s}}{(9.11 \times 10^{-31} \text{ kg})(3.0 \times 10^7 \text{ m/s})}$$

$$\lambda = 2.4 \times 10^{-11} \text{ m} = 2.4 \times 10^{-2} \text{ nm}$$

The mass must be in kg for units to cancel properly because $1 \text{ J} = 1 \text{ kg m}^2/\text{s}^2$.

(b) Convert 55 g to 0.055 kg, then use the same equation.

$$\lambda = \frac{6.63 \times 10^{-34} \text{ J s}}{(0.055 \text{ kg})(35 \text{ m/s})} = 3.4 \times 10^{-34} \text{ m}$$

$$\lambda = 3.4 \times 10^{-25} \text{ nm}$$

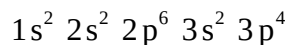
This wavelength is so small that it is insignificant on the scale of the ball. We cannot detect a wavelength this small in this situation; the wave properties of such large objects do not affect ordinary observations.

Question 28 Solution

- (a) False. It takes less energy to ionize the electron from $n = 3$ than from the ground state.
- (b) True.
- (c) False. The energy difference between $n = 3$ and $n = 2$ is smaller than that between $n = 3$ and $n = 1$. The wavelength is therefore longer for 3-to-2 than for 3-to-1. Energy and wavelength are inversely proportional: $E = hc/\lambda$.
- (d) True. The energy difference between the same two levels has the same magnitude in either direction.
- (e) False. The $n = 2$ level is the first excited state and $n = 3$ is the second excited state.

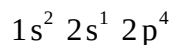
Question 29 Solution

- (a) The complete ground-state configuration is:



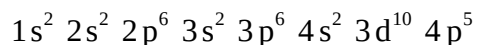
There are $2 + 2 + 6 + 2 + 4 = 16$ electrons. The atom is neutral, so it also has 16 protons: sulfur, S.

- (b) The complete excited-state configuration is:



The neutral atom has $2 + 1 + 4 = 7$ electrons and therefore seven protons: nitrogen, N.

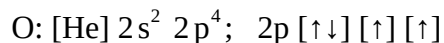
- (c) The complete ground-state configuration is:



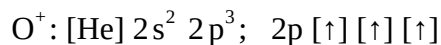
The ion has 35 electrons. Since its charge is -1 , it has 34 protons, identifying selenium. The ion is Se^- .

Question 30 Solution

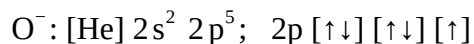
Find the unpaired electrons by examining the incompletely filled subshells. Each bracketed group represents an orbital; opposite arrows represent opposite spins.



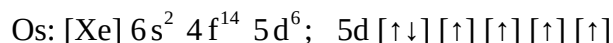
Two unpaired electrons.



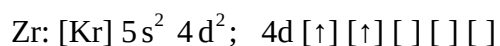
Three unpaired electrons.



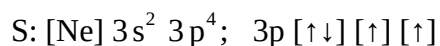
One unpaired electron.



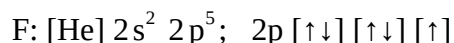
Four unpaired electrons.



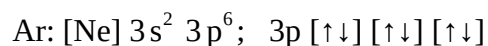
Two unpaired electrons.



Two unpaired electrons.



One unpaired electron.



Zero unpaired electrons.

Question 31 Solution

Atomic radius generally decreases from left to right across a period and increases from top to bottom in a group.

(a) $\text{S} < \text{Se} < \text{Te}$

(b) $\text{Br} < \text{Ni} < \text{K}$

(c) $\text{F} < \text{Si} < \text{Ba}$

All three orders follow the general radius trend.

Question 32 Solution

The general first-ionization-energy trend is opposite to the radius trend: it increases from left to right across a period and decreases from top to bottom in a group.

(a) $\text{Te} < \text{Se} < \text{S}$

(b) $\text{K} < \text{Ni} < \text{Br}$

(c) $\text{Ba} < \text{Si} < \text{F}$

All three orders follow the general first-ionization-energy trend.