SCH4U Chem 12 Chapter 3

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1. (a) Thomson's model of the atom focussed on the electrons. Rutherford's model focussed on the nucleus. Thomson viewed the atom as a positively charged sphere, embedded with sufficient electrons to balance the charge. Rutherford pictured the atom as mostly empty space, with electrons orbiting around a massive region of concentrated positive charge at the centre (the nucleus).

(b) Rutherford based his model on evidence from his alpha particle scattering experiments. In these experiments, most alpha particles passed through the gold foil with almost no deflection. About one in every 8000, however, was deflected significantly. A few of the particles even bounced back. These results were inconsistent with Thomson's model, which predicted that alpha particles would be deflected by no more than 0.005°.

2. There were two concerns with Rutherford's model:

- If all the positive charges were concentrated in a very tiny nucleus, they should fly apart, since like charges repel.
- Rutherford's model contradicted the assumptions of nineteenth-century physics. According to these assumptions, an electron circling around a nucleus would continually give off energy and would eventually collapse into the nucleus. However, most atoms are stable. Also, the continuous emission of energy would give a complete spectrum of colours, but hydrogen does not.

3. (a) The model I have created for the atom differs from the model developed by Rutherford. In my model, the electrons can only have certain fixed energy values. The energy of an electron is thus quantized, and an electron does not emit energy continuously. Energy is only absorbed or emitted when an electron moves from one energy level to another. The discrete packages of light that are absorbed or emitted are called photons. Using my model, the hydrogen spectrum can be explained as follows:

The gas discharge tube provides the quantum of energy necessary to promote hydrogen's electron from one energy level to a higher energy level, which is farther from the nucleus. As the electron drops back from the higher energy level, a photon of light is released and a thin, coloured line appears in the emission spectrum.

(b) Mr Bohr, your ideas about electrons are contrary to the laws of physics. Electrons moving around a positive nucleus will continuously emit energy, like the energy emitted by the sun. Thus, your model, like Rutherford's, is flawed. Also, the idea of discrete packages of light, which you refer to as photons, is an absurd idea. Light and other electromagnetic radiation travel as waves, not as particles. What evidence do you have to support your ideas about electrons?

4. (a) The types of electromagnetic radiation, from longest to shortest wavelengths, are radio waves, microwaves, infrared radiation, visible light, ultraviolet radiation, X-rays, and gamma rays.

(b) The types of electromagnetic radiation, from highest to lowest frequencies, are gamma rays, X-rays, ultraviolet radiation, visible light, infrared radiation, microwaves, and radio waves.

5. Visible light radiation and ultraviolet radiation both travel at the speed of light. They both have wave-like and particle-like properties. However, ultraviolet radiation is not visible, and it has a shorter wavelength and higher frequency than visible light radiation.

6. (a) As an electron moves from $n = 1$ to $n = 6$, it is moving from a lower to a higher energy level. Therefore, the electron is absorbing a photon of energy.

(b) As an electron is moving from $n = 5$ to $n = 2$, it is moving from a higher to a lower energy level. Therefore, the electron is emitting a photon of energy.

(c) As an electron is moving from $n = 4$ to $n = 3$, it is moving from a higher to a lower energy level. Therefore, the electron is emitting a photon of energy.

7. The quantum number n determines the location of the electron orbit relative to the nucleus. The higher the value of n, the farther from the nucleus the orbit is, and the more energy that is required to move the electron there. Therefore, the order of energy from highest to lowest is $n = 7$, $n = 5$, $n = 4$, $n = 2$, and $n = 1$.

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1. Bohr, like Thomson and Rutherford, visualized the electron as a particle. The term "orbits" and the phrase "jump from one energy level to another" are particle-based. The quantum mechanical model of the atom is based on an electron having wave properties. It describes the motion of an electron using the mathematics of waves.

2.

Quantum Number	Symbol	Property of Orbital Described
principal	n	specifies the energy level and relative size of an orbital
orbital-shape	l	refers to the energy sublevels within the main energy level and determines the shape of the orbital
magnetic	m_l	describes the orientation of the orbital around the nucleus

3.

n	l	m_l
4	0	0
	1	−1, 0, +1
	2	−2, −1, 0, +1, +2
	3	−3, −2, −1, 0, +1, +2, +3

4. Points of agreement:
   - In both models, value of n is interpreted as the relative energy and size.
   - Changes of n involve the absorption or emission of a quantum of energy.
   - n has values of $n = 1, 2, 3$, and so on.

   Point of disagreement:
   - In the Bohr model, n refers to the orbit of an electron. In the quantum mechanical model, n refers to the energy and size of an orbital.

5. **(a)** The value of l and the name $1p$ are incorrect. When $n = 1$, the only value of l is 0. The first energy level has no sublevels, so $1p$ is impossible.

   (b) If the name is $4d$, the value for l should be 2, not 3. For $l = 3$, the type of orbital is f and the name should be $4f$.

   (c) If the name is $3p$, the value of m_l is incorrect. When $l = 1$, the acceptable values of m_l are 0, −1, and +1.

6. The missing values are boldfaced.

   (a) $n = \mathbf{4}$, $l = \mathbf{1}$, $m_l = 0$; name: $4p$

   (b) $n = 2$, $l = 1$, $m_l = 0$; name: $\mathbf{2p}$

   (c) $n = 3$, $l = 2$, $m_l = -2$; name: $\mathbf{3d}$

   (d) $n = \mathbf{2}$, $l = \mathbf{0}$, $m_l = \mathbf{0}$ name: $2s$

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1. (a) The fourth quantum number is the spin quantum number. It is different from the other three quantum numbers because it is not part of the solutions to the wave equations that are used to describe electrons. It comes from empirical evidence about the electron.

   (b) The fourth quantum number has only two values: $+\frac{1}{2}$ and $-\frac{1}{2}$.

2. The orbitals of a hydrogen atom, like the orbitals of other atoms, are mathematical solutions to the wave equations that are used to describe electrons. Each orbital can be characterized by three quantum numbers—n, l, and m_l—and can have a maximum of only two electrons. However, for multi-electron atoms, orbitals of different types in the same energy level (same n) have different energy values. For example, the $2p$ orbital always has slightly higher energy value than the $2s$ orbital, whereas the $2s$ and $2p$ orbitals for hydrogen have the same energy value.

3.

	Electron Configuration	Group	Period	Orbital Block
(a)	[Kr] $5s^2\ 4d^1$	3 (IIIB)	5	*d* block
(b)	[Ar] $4s^2\ 3d^{10}\ 4p^3$	15 (VA)	4	*p* block
(c)	[He] $2s^2\ 2p^6$	18 (VIIIA)	2	*p* block
(d)	[Ne] $3s^2\ 3p^1$	13 (IIIA)	3	*p* block

4. Since a valence electron is always in a higher energy level than an electron in an inner energy level, less energy is needed to remove the valence electron. An electron in a higher energy level is farther away from the nucleus. Thus, the attractive force between the nucleus and the electron is less, and the effective charge is greater. It is more easily removed from its atom.

5. An *s* block element from groups 1 (IA) and 2 (IIA), is most likely to form a cation when a valence electron is removed from its atom. A cation is a positively charged ion. Thus, electrons must be lost from a neutral atom to form a cation. The atomic property of atoms that is related to the formation of a cation is the ionization energy. The *s* block elements have the lowest first ionization energies within the same period.

6. (a) The aufbau principle, the exclusion principle, and Hund's rule are all used for drawing orbital diagrams of atoms. The aufbau principle gives the general order for filling of orbitals with different energy values. The exclusion principle puts a limit of two electrons in an orbital. Hund's rule states that electrons add singly to separate orbitals of the same energy before they pair up.

   (b) Electron configurations are shorthand notations for orbital diagrams.

   (c) In general, as the atomic radius increases, the ionization energy decreases. Similarly, as the atomic radius decreases, the ionization energy increases.

   (d) Ionization energy is the quantity of energy required to remove an electron from a gaseous atom, whereas electron affinity is the quantity of energy released when an electron is added to a gaseous atom. Ionization energy involves the formation of a cation, and electron affinity involves the formation of an anion.

   (e) Valence electrons experience a smaller Z_{eff} than inner electrons because the inner electrons shield or screen the valence electrons from the attractive force of the nucleus.

   (f) The periodic table organizes elements so that elements in each vertical column have similar chemical properties.

   (g) The period of repetition in the periodic table is related to the types of orbitals that are being filled. The electron configurations for elements provide this information. As well, the periodic table can also be used to determine electron configurations.

   (h) Electron configurations can be used to deduce the set of quantum numbers that represents each electron in the atom. For example, the electron configuration for nitrogen, N, is [He] $2s^2\ 2p^3$. The three electrons in the $2p$ orbital have the following sets of four quantum numbers:

   $n = 2,\ l = 1,\ m_l = -1,\ m_s = +\frac{1}{2}$

   $n = 2,\ l = 1,\ m_l = 0,\ m_s = +\frac{1}{2}$

   $n = 2,\ l = 1,\ m_l = +1,\ m_s = +\frac{1}{2}$

   (i) The largest number before an *s* orbital in the electron configuration gives the period number. The number of electrons in the last part of the condensed electron configuration gives the group number. For example, the electron configuration for nitrogen, N, is [He] $2s^2\ 2p^3$. Nitrogen is in period 2 and is five elements away from He, so it is in group 15 (VA).

   (j) The quantum mechanical model of the atom provides electron configurations that explains the periodic repetition of chemical properties.

7. (a) Na^+: $1s^2\ 2s^2\ 2p^6$ or [Ne]

   (b) Ca^{2+}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ or [Ar]

   (c) Cl^-: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ or [Ne] $3s^2\ 3p^6$ or [Ar]

   (d) S^{2-}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ or [Ne] $3s^2\ 3p^6$ or [Ar]

   These elements tend to gain or lose electrons to form ions with the electron configurations of the nearest noble gas.

8. The largest jump in ionization energy is between the second and third ionization energies. The first two electrons are likely in the same energy level, and the third electron is likely in an energy level that is closer to the nucleus. The atom appears to have two valence electrons that are probably *s* type electrons, because the large difference between the second electron and third electron could not be due to the change from one type of orbital to another. The large difference must be due to a change of energy levels. Thus, the predicted valence electron configuration for this atom is ns^2.

Answers to Knowledge/ Understanding Questions

1. Rutherford noticed, in his gold foil experiment, that most of the alpha particles passed through the gold foil with almost no deflection. About one in every 8000, however, bounceed back (scattered). Based on this evidence, Rutherford proposed that the atom is mostly empty space. The bouncing back of the alpha particles suggested that most of the mass of the atom is in a very tiny and dense nucleus of the atom.
2. Rutherford's nuclear model successfully accounted for the observations of alpha particle scattering, but it raised questions. If the electron were circling the nucleus, then classical physics predicted that it would give off energy, resulting in the electron spiralling into the nucleus. However, atoms did not collapse, and neither did they give off a continuous spectrum. For example, the atomic spectrum of hydrogen appeared as lines instead of a continuous smear of colour, like the spectrum of the sun. These observations and inferences led Bohr to place restrictions on electrons. These restrictions allowed him to account for the atomic spectra of hydrogen. He and Max Planck introduced the idea that the electrons in an atom can only have certain energies: in other words, the energy of an electron in an atom is quantized.
3. Both Rutherford and Bohr proposed atomic models that had a nucleus, with electrons orbiting around the nucleus. In Bohr's model, the paths of the electrons were fixed—they could only be at certain distances from the nucleus. In Rutherford's model, however, electrons were free to spin around the nucleus at different distances from the nucleus. There were no fixed paths. Bohr's description of an atom was closer to a planetary model than Rutherford's, since Earth does follow a predictable path around the sun.
4. **(a)** Planck proposed that matter, at the atomic level, could absorb or emit only discrete quantities of energy. The energy of an atom was quantized.
   (b) De Broglie proposed that matter had wave-like properties.
   (c) Einstein proposed that light was also quantized and had particle-like properties. The quanta of light were later called photons.
   (d) Heisenberg showed that it was not possible to know both a particle's position and its momentum precisely. He and Schrödinger are credited with introducing quantum mechanics.
   (e) Schrödinger used mathematics and statistics to combine de Broglie's idea of matter waves and Einstein's idea of quantized energy particles (photons). Schrödinger's ideas, together with the Heisenberg's uncertainty principle, led to a description of atomic particles in terms of wave equations.
5. The *s* block includes hydrogen, helium, and the elements of group 1 (IA) and group 2 (IIA). Valence electrons occupy only the *ns* orbital in this block. The *p* block includes elements from group 13 (IIIA) to group 18 (VIIIA). Electron configurations of the *p* block atoms take the general form ns^2np^a, where *a* represents a number from 1 to 6. The *d* block includes all the transition elements from group 3 (IIIB) to group 12 (IIB). In general, *d* block elements have filled *ns* orbitals, as well as filled or partially filled *d* orbitals. The *f* block includes all the inner transition elements. Atoms of *f* block elements have filled orbitals in the outer energy levels, as well as filled or partially filled 4*f* or 5*f* orbitals.
6. Pauli exclusion principle restricted the number of electrons in an orbital to be no more than two. Thus, the maximum number of electrons is 2 for an *s* type orbital, 6 for a *p* type, 10 for a *d* type, and 14 for an *f* type. According to Hund's rule, when electrons are added to a *p*, *d*, or *f* type orbital, electrons of the same spin are added separately to orbitals with the same energy first, before any pairing of electrons occurs. Thus, when writing the electron configuration of the first three electrons with $l = 1$ and $m_l = 0$, $+1$ and -1, all three electrons have $m_s = +\frac{1}{2}$.
7. The electron is in the *p* orbital ($l = 1$) of the third energy level ($n = 3$): that is, the 3*p* orbital.
8. An excited state sulfur atom is a sulfur atom that has absorbed a quantity of energy so that one or more of its ground state electrons are moved to an orbital of higher energy. The ground state electron configuration for sulfur is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$. Examples of excited states can be $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3\ 4s^1$ and $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3\ 5s^1$.
9. The second member of each of the following pairs has a larger radius because it is farther down a group in the periodic table. Each period of the period table begins with the first member of a new and higher energy level, resulting in a larger radius of the atom.
   (c) Ca, Sr
   (d) Al, Ga
   (f) Cl, Br

   For the two pairs in **(a)** and **(e)**, Na and O have the larger atomic radius since atomic radius decreases across a period. The effective nuclear charge increases across a period, pulling the electrons closer to the nucleus. For **(b)**, K has the lager atomic radius because Cl is near the end of period 3 and K is at the beginning of the next period, where a new and higher energy level begins.
10. U, *f* block, inner transition; Zr, *d* block, transition; Se, *p* block, main group; Rb, *s* block, main group; Re, *d* block, transition; Sr, *s* block, main group; Dy, *f* block, inner transition; Kr, *p* block, main group

11. The formulas for the oxides are Na_2O, K_2O, MgO, CaO, Al_2O_3, and Ga_2O_3. Each member of a group has the same outermost, or highest energy, electrons. Knowing the formula of one member of a group enables you to know the formulas of the other members of the same group.

12. **(a)** The ground state electron configuration for arsenic, As, is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^3$. Each filled orbital should contain two electrons. According to Hund's rule, each of the three $4p$ orbitals has one un-paired electron. The total number of orbitals is 18.

(b) Other than the three $4p$ orbitals that are half filled, 15 of the orbitals are completely filled.

(c) There are four orbitals associated with the fourth energy level: one $4s$ and three $4p$.

(d) Theoretically speaking, one electron can occupy one orbital in the excited state. The maximum number would be 33 and the minimum is 17.

Answers to Inquiry Questions

13. The electron configuration for vanadium, V, is [Ar] $4s^2\ 3d^3$. Therefore, diagram **(c)** would be the correct orbital diagram. According to the aufbau principle, the $4s$ orbital fills first, before the $3d$ orbital. When the $4s$ orbital is filled, the three electrons fill the $3d$ orbitals following Hund's rule. That is, the three electrons of the same spin are added separately to the three $3d$ orbitals before any pairing of electrons occurs.

14. **(a)** The electron configuration for carbon is $1s^2\ 2s^2\ 2p^2$. According to Hund's rule, the two $2p$ electrons both have the same spin and should be in two different orbitals. Thus, in the diagram, the last two arrows should point upward and there should only be one arrow in each of the first two boxes in the last row of $2p$ orbitals.

(b) The electron configuration for iron is [Ar] $3d^6\ 4s^2$. Since the $3d$ orbitals fill before the $4p$ orbitals, the arrow in one of the $4p$ orbitals should point downward and move to pair up with the $3d$ electron in the first box of the row of five $3d$ orbitals.

(c) The electron configuration for bromine is [Ar] $3d^{10}\ 4s^2\ 4p^5$. Thus, two arrows pointing downward should be added to the first two boxes of the $4p$ orbitals to pair up the electrons.

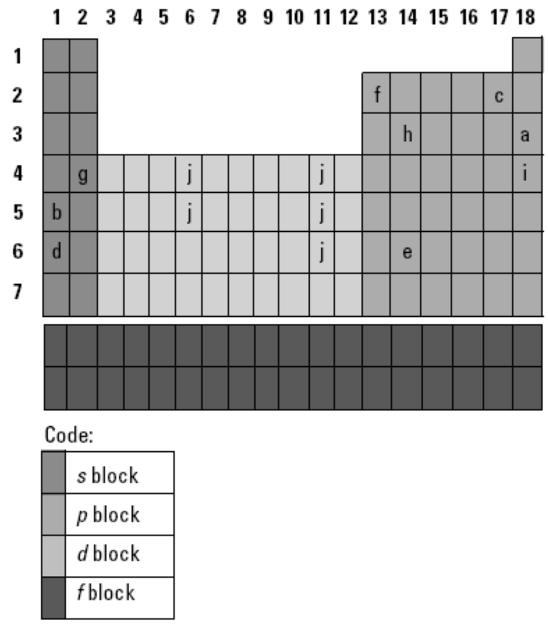
15. **(a)** magnesium: ground state electron configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2$

(b) chlorine: ground state electron configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

(c) manganese: ground state electron configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^5$

(d) yttrium: ground state electron configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^1$

16. and 17.



18. For element 126, the ground state electron configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^6\ 7s^2\ 5f^{14}\ 6d^{10}\ 7p^6\ 8s^2\ 5g^6$ or [Rn] $7s^2\ 5f^{14}\ 6d^{10}\ 7p^6\ 8s^2\ 5g^6$. Thus, the outermost electron is in $8s^2$, so $n = 8$ and $l = 0$. There would be just one s orbital in this energy sublevel.

Answers to Communication Questions

19. **(a)** K is in group 1 (IA), period 4, and Sr is in group 2 (IIA), period 5. The atomic size of K is smaller than the atomic size of Rb, which is larger than the atomic size of Sr. Between K and Rb, an additional energy level makes Rb larger. An increase in the effective charge makes Sr smaller than Rb. However, neither effect gives a quantitative measure, so it is hard to know whether K or Sr has a larger atomic size. In actual fact, the atomic size of K is 227 pm, of Rb is 248 pm, and of Sr is 215 pm. Thus, the effective charge is the more important factor here.

(b) The electron configuration for Mn is [Ar] $3d^5\ 4s^2$. The electron configuration for Fe is [Ar] $3d^6\ 4s^2$. Both Mn and Fe have two electrons in the $4s$ orbital. Ranking these two elements according to their first ionization energies could be tricky because of the half-filled d type orbital. The atomic radius of Mn is 127 pm and the atomic radius of Fe is 126 pm, as expected by the increase in the effective charge. Thus, the first ionization energy of Fe is likely to be higher than the first ionization energy of Mn. From the graph of first ionization energies, Fe has a higher first ionization energy than Mn. Apparently, the half-filled *d* type orbital does not alter the trend for these elements.

(c) Na is in group 1 (IA), period 3, and Ca is in group 2 (IIA), period 4. The first ionization energy generally decreases down a group and increases across a period. Within the same group, K has a lower first ionization energy than Na, since K is in the next higher energy level. Also, across the same period, K has a lower first ionization energy than Ca. Again, a quantitative measure is needed to rank the ionization energies, and thus the metallic character, of Na and Ca.

20. Atoms are not solid spheres, and their volume is described in terms of probabilities. Therefore, the size of an atom has to be measured indirectly by first measuring the distance between the nuclei of the bonded, neighbouring atoms. The atomic radius is then calculated from the bond length.

21. (a) For valence electron configuration $5s^1$, the period number of the element comes from the 5 of the $5s$ orbital. The position in the period is determined by the total number of valence electrons, which is 1. The element is the first element of period 5, Rb.

(b) For valence electron configuration $4s^2\ 3d^2$, the period is determined by the 4 in the $4s$ orbital and the position in the period by the total number of valence electrons, which is 4. Thus, the element is in the fourth element of period 4, Ti.

(c) For valence electron configuration $3s^2$, the element is the second element in period 3, Mg.

(d) For valence electron configuration $4s^2\ 3d^{10}\ 4p^3$, the element is the fifteenth element in period 4, As.

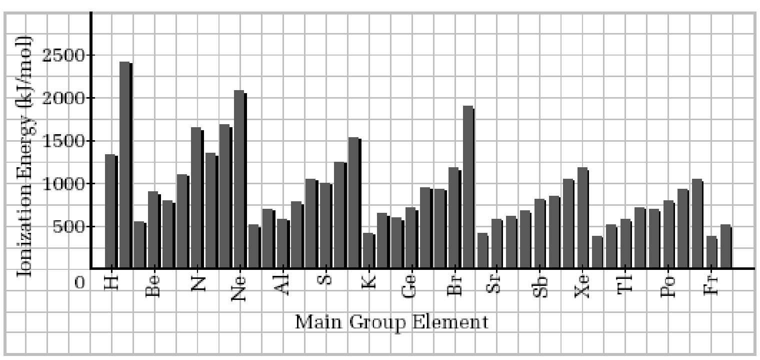
22. When the wave equations were solved for an electron, the rules for the first three quantum numbers were determined: $n = 1, 2, 3 \ldots n$, and for every value of n, $l = 0, \pm 1, \pm 2, \pm 3 \ldots \pm(n - 1)$, and for every value of l, $m_l = 0, \pm 1, \pm 2, \pm 3 \ldots \pm l$. When $n = 1$, l can only have one value, 0. Thus, there is only one type of orbital, *s* orbital, in the first energy level (period 1).

23. (a) The ground state electron configuration for scandium is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^1$. The ground state electron configuration for aluminum is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$.

(b) Since there are three outermost, or valence, electrons, scandium and aluminium might be expected to form 3+ ions. Since the composition of the oxides of both elements implied a 3+ ion, this would further encourage chemists to group them together.

(c) The placement of scandium and aluminum in different groups is based on trends in properties (such as atomic radii and ionization energies) and on electron configurations. For example, if scandium was between aluminum and gallium in group 13 (IIIA), then the trend in atomic radius and ionization energy would have a discontinuity at scandium. The radius would gradually increase from boron to aluminum, and then take a big increase in size to scandium and a big decrease in size to gallium. There is a similar discontinuity in ionization energies. Thus, scandium fits in with the trends better where it is.

24.



Answers to Making Connections Questions

25. There are no correct or typical answers to questions like this, but there are some things to look for in the argument presented by students.

The first is a list of applications of plutonium. Plutonium is used for

- nuclear weapons
- a source of nuclear power
- pacemakers

Students should then list properties that make plutonium useful in these applications:

- The most important isotope of plutonium is plutonium-239, which has a half-life of 24 360 years.
- Plutonium-239 readily undergoes fission. It can be both used and produced in quantity in nuclear reactors. In addition to being fissionable, it can be split by neutrons of very low (ideally zero) energy. Thus, it can be assembled into a critical mass, and hence can sustain a chain reaction without an external source of neutrons. This makes it useful for nuclear weapons.
- Plutonium-238 has been used to power pacemakers and equipment on the moon, using the heat it emits.

Next, students should discuss the risks and the benefits:

- The risk of working with plutonium is the high radioactivity. The benefit is the energy produced.
- Plutonium-239 is an extremely hazardous poison due to its high radioactivity.
- To reduce the risks, companies need to invest in costly protection for workers and the environment. Since radioactivity is very penetrating, thick concrete walls, huge vats of heavy water, and massive steel containers are required to contain it. Therefore, building a reactor and its containment systems is very expensive. Dismantling the facility once it can no longer be used is also expensive, since much of the waste material is radioactive.
- Economists say that power plants are cheaper to operate than coal-fired or natural-gas-fired electric generating plants. However, when higher construction costs for new reactors and the price of decommissioning old plants and disposing of waste are factored in, coal and natural gas plants are cheaper overall.

Finally, students have to justify their opinions on whether the benefits outweigh the risks.

- Their opinions will be very personal and not likely to be based on evidence alone. There is much material, both for and against the use of nuclear energy. Students may have difficulty sorting through all the arguments regarding the costs and the alternatives, but the issue is important and worth the effort.

26. Again there are no right or wrong answers to this question. Because of the enormity of the issue, students should focus on the research conducted by one institution. They might first look at the mission statement and goals of the institution and then justify if the money should be spent elsewhere.

Chapter 3

Atoms, Electrons, and Periodic Trends

Solutions for Practice Problems

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1. **Problem**

   What are the allowed values for l in each of the following cases?

   (a) $n = 5$

   (b) $n = 1$

   Solution

   For any value of the quantum number n, the allowed values of the quantum number l are integers ranging from 0 to $(n - 1)$.

   (a) $n = 5$, $\therefore (n - 1) = 4$, and l can have values 0, 1, 2, 3, or 4.

   (b) $n = 1$, $\therefore (n - 1) = 0$, and l can have only the value 0.

   Check Your Solution

   For any value of n, there are n possible values of l. When $n = 5$, there were five allowed values of l. When $n = 1$, there was one allowed value for l. The answer is correct.

2. **Problem**

   What are the allowed values for m_l for an electron with the following quantum numbers?

   (a) $l = 4$

   (b) $l = 0$

   Solution

   For any value of the quantum number l, the allowed values of the quantum number m_l are integers ranging from $-l \ldots 0 \ldots +l$.

   (a) For $l = 4$, m_l can have the values –4, –3, –2, –1, 0, +1, +2, +3, and +4.

   (b) For $l = 0$, m_l can only have the value 0.

   Check Your Solution

   For any value of l there can be $(2l + 1)$ values of m_l. For $l = 4$, there are $2(4) + 1 = 9$ values for m_l. For $l = 0$, there is $2(0) + 1 = 1$ value for m_l.

3. **Problem**

   What are the names, m_l values, and total number of orbitals described by the following quantum numbers?

   (a) $n = 2$, $l = 0$

   (b) $n = 4$, $l = 3$

Solution
The type of orbital is determined by the value of n and l as $n(l)$. The orbital is s, for $l = 0$, p for $l = 1$, d for $l = 2$ and f for $l = 3$. You can find the possible values for m_l from l. The total number of orbitals is given by the total number of m_l values.

(a) For $l = 0$, m_l can have only the value 0. Since $n = 2$ and $l = 0$, the quantum numbers represent a $2s$ orbital. Since there is only one allowed value for m_l, there is one $2s$ orbital.

(b) For $l = 3$, m_l can have allowed values of –3, –2, –1, 0, +1, +2, and +3. Since $n = 4$ and $l = 3$, the quantum numbers represent a $4f$ orbital. Since there are seven allowed values for m_l, there are seven $4f$ orbitals.

Check Your Solution
Since the total number of orbitals for any given l value is $(2l + 1)$, there should be $2(0) + 1 = 1$ orbital when $l = 0$, and $2(3) + 1 = 7$ orbitals when $l = 3$. This answer is correct.

4. Problem
Determine the n, l, and possible m_l values for an electron in the $2p$ orbital.

Solution
The type of orbital takes its name from the value of the quantum numbers n and l. A p orbital corresponds to $l = 1$. Since it is a $2p$ orbital, $n = 2$. You can find the possible values for m_l from l. For $l = 1$, the allowed values that are possible for m_l are –1, 0, and +1.

Check Your Solution
For any value of l, there can be $(2l + 1)$ values of m_l. For $l = 1$, there are $2(1) + 1 = 3$ values for m_l. This answer is correct.

5. Problem
Which of the following are allowable sets of quantum numbers for an atomic orbital? Explain your answer in each case.

(a) $n = 4$, $l = 4$, $m_l = 0$
(b) $n = 3$, $l = 2$, $m_l = 1$
(c) $n = 2$, $l = 0$, $m_l = 0$
(d) $n = 5$, $l = 3$, $m_l = -4$

Solution
For any value of the quantum number n, the allowed values of the quantum number l are integers ranging from 0 to $(n - 1)$. For any value of the quantum number l, the allowed values of the quantum number m_l are integers ranging from $-l \ldots 0 \ldots +l$. Apply these criteria to each case.

(a) For $n = 4$, the allowed values for l are 0, 1, 2, or 3. $l = 4$ is not an allowed value. It is allowable for m_l to have the value 0. This combination of quantum numbers is not allowable.

(b) For $n = 3$, the allowed values for l are 0, 1, or 2. For the value of $l = 1$, m_l can have the value 1. This combination of quantum numbers is allowable.

(c) For $n = 2$, the allowed values for l are 0 or 1. For the value of $l = 0$, m_l can have the value 0. This combination of quantum numbers is allowable.

(d) For $n = 5$, the allowed values for l are 0, 1, 2, 3, or 4. For $l = 3$, m_l cannot have the value –4. This combination of quantum numbers is not allowable.

Check Your Solution
In (b) and (c), the criteria for allowed quantum numbers are met.

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6. **Problem**
   Use the aufbau principle to write complete electron configurations and complete orbital diagrams for atoms of the following elements: sodium, magnesium, aluminum, silicon, phosphorus, sulfur, chlorine, and argon (atomic numbers 11 through 18).

   Solution
   (See solution for Problem 7)

7. **Problem**
   Write condensed electron configurations for atoms of these same elements.

   Solution (Problem 6 and Problem 7)
   The aufbau principle is a process of building up the ground state electron configuration of atoms, by adding one electron to the lowest available energy level in order of increasing atomic number. To build up to element 18, the order in which orbitals are filled increases: $1s<2s<2p<3s<3p$. The Pauli Exclusion Principle specifies that no two electrons in an atom can have the same set of four quantum numbers. This leads to the conclusion that an *s* orbital can hold a maximum of two electrons and the three *p* orbitals at any energy level can hold, in total, a maximum of six electrons. Electrons in the same orbital are shown to have opposite spin, ↑↓, a consequence of the Pauli Exclusion Principle that is referred to as Hund's Rule. A condensed electron configuration represents the electrons in the filled inner core by the symbol of the noble gas to which it corresponds, followed by the electron configuration of the valence shell.

Element	Z	Electron configuration	Orbital diagram 1*s*	2*s*	2*p*	3*s*	3*p*	Condensed configuration
Na	11	$1s^22s^22p^63s^1$	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑		$[Ne]3s^1$
Mg	12	$1s^22s^22p^63s^2$	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓		$[Ne]3s^2$
Al	13	$1s^22s^22p^63s^23p^1$	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑	$[Ne]3s^23p^1$
Si	14	$1s^22s^22p^63s^23p^2$	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑ ↑	$[Ne]3s^23p^2$
P	15	$1s^22s^22p^63s^23p^3$	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑ ↑ ↑	$[Ne]3s^23p^3$
S	16	$1s^22s^22p^63s^23p^4$	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑ ↑	$[Ne]3s^23p^4$
Cl	17	$1s^22s^22p^63s^23p^5$	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑↓ ↑	$[Ne]3s^23p^5$
Ar	18	$1s^22s^22p^63s^23p^6$	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑↓ ↑↓	$[Ne]3s^23p^6$

   Check Your Solution
   The number of electrons increases in order of increasing atomic number, and the sequence predicted by the aufbau principle is followed. In any orbital containing two electrons, the electrons are of opposite spin.

8. **Problem**
   Make a rough sketch of the periodic table for elements 1 through 18, including the following information: group number, period number, atomic number, atomic symbol, and condensed electron configuration.

Solution

	1 IA	2 IIA	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
1	1 H 1.01 $1s^1$							2 He 4.003 $1s^2$
2	3 Li 6.941 $[He]2s^1$	4 Be 9.012 $[He]2s^2$	5 B 10.81 $[He]2s^22p^1$	6 C 12.01 $[He]2s^22p^2$	7 N 14.01 $[He]2s^22p^3$	8 O 16.00 $[He]2s^22p^4$	9 F 19.00 $[He]2s^22p^5$	10 Ne 20.18 $[He]2s^22p^6$
3	11 Na 22.99 $[Ne]3s^1$	12 Mg 24.13 $[Ne]3s^2$	13 Al 26.98 $[Ne]3s^23p^1$	14 Si 28.09 $[Ne]3s^23p^2$	15 P 30.97 $[Ne]3s^23p^3$	16 S 32.07 $[Ne]3s^23p^4$	17 Cl 35.45 $[Ne]3s^23p^5$	18 Ar 39.95 $[Ne]3s^23p^6$

Check Your Solution
The arrangement of elements 1–18 follows a regular pattern. The number of energy levels in any horizontal row corresponds to the period number, and the number of electrons in the valence shell corresponds to the group number (or to the last digit of the group number).

9. Problem
A general electron configuration for atoms belonging to any element of group 1 (IA) is ns^1, where n is the quantum number for the outermost occupied energy level. Based on the patterns you can observe so far for elements 1 to 18, predict the general electron configuration for the outermost occupied energy levels of groups 2 (IIA), 13 (IIIA), 14 (IVA), 15 (VA), 16 (VIA), 17 (VIIA), and 18 (VIIIA).

Solution
Referring to the solution given for problem 8, it can be observed that within any group, the electrons in the valence shell are found in the same type of orbitals, and the number of valence electrons is equal to the group number or last digit of the group number. Using these criteria, the pattern should be the same for all elements within the group: 2 (IIA): ns^2; 13 (IIIA): ns^2np^1; 14 (IVA): ns^2np^2; 15 (VA): ns^2np^3; 16 (VIA): ns^2np^4; 17 (VIIA): ns^2np^5; 18 (VIIIA): ns^2np^6.

Check Your Solution
The number of the electrons in the valence shell of elements within a group follows the expected pattern — the number of electrons is equal to the number of the group or the last digit of the group number. The link between the orbital notation and the group number has been applied correctly.

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10. Problem
Without looking at a periodic table, identify the group number, period number, and block of an atom that has the following electron configuration.

(a) $[Ne]3s^1$
(b) $[He]2s^2$
(c) $[Kr]5s^24d^{10}5p^5$

Solution

(a) This element corresponds to the general notation ns^1, which is an *s* block element, since there are only *s* electrons in the valence shell. $n = 3$, so the element is in period 3. Since there is only one valence electron, the element is in group 1.

(b) This element corresponds to the general notation ns^2, which is an *s* block element, since there are only *s* electrons in the valence shell. $n = 2$, so the element is in period 2. Since there are two valence electrons, the element is in group 2.

(c) This element corresponds to the general notation ns^2np^5, which is a *p* block element, since there are *s* and *p* electrons in the valence shell. $n = 5$, so the element is in period 5. Since there are seven valence electrons, the element is in group 17.

Check Your Solution

The link between *n* value and period, and the match between the valence electrons for each element and the general notation for elements in that group have been correctly applied.

11. Problem

Use the aufbau principle to write the complete electron configurations for the atom of the element that fits the following descriptions:

(a) group 2 (IIA) element in period 4

(b) noble gas in period 6

(c) group 12 (IIB) element in period 4

(d) group 16 (IVB) element in period 2

Solution

(a) Group 2 elements have the general notation ns^2, and for an element in period 4, $n = 4$. Therefore, the valence electrons are represented by $4s^2$. Also, an element in period 4 will have an inner core that corresponds to the noble gas in period 3 of the periodic table, which is argon. The element is Ca, and its complete electron configuration is $1s^22s^22p^63s^23p^64s^2$.

(b) The noble gas in period 6 will have 6 energy levels and have the general notation $6s^26p^6$. Follow the sequence outlined in the aufbau principle until you reach $6p^6$. This should be the noble gas radon. Its complete electron configuration is $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^{10}6p^6$

(c) Group 12 elements have the general notation $ns^2(n-1)d^{10}$, and for an element in period 4, $n = 4$. Also, an element in period 4 will have an inner core that corresponds to the noble gas in period 3 of the periodic table, which is argon. The condensed electronic configuration will be $[Ar]4s^23d^{10}$. The element is Zn, and its complete electron configuration is $1s^22s^22p^63s^23p^64s^23d^{10}$.

(d) Group 16 elements have the general notation ns^2p^4, and for an element in period 2, $n = 2$. Also, an element in period 2 will have an inner core that corresponds to the noble gas in period 1 of the periodic table, which is helium. The element is O, and its complete electron configuration is $1s^22s^22p^4$.

Check Your Solution

A check of the complete electron configurations shows that the link between *n* and period number, the sequence predicted by the aufbau principle, and the match between the valence electrons for each element and the general notation for elements in that group are correctly followed.

12. Problem

Identify all the possible elements that have the following valence electron configurations.

(a) s^2d^1
(b) s^2p^3
(c) s^2p^6

Solution

(a) s^2d^1 is a *d* block configuration that corresponds to the general notation $ns(n-1)d$. Elements having this configuration are in group 3, and include Sc, Y, La, and Ac.
(b) s^2p^3 is a *p* block configuration that corresponds to the general notation ns^2np^3. Elements having this configuration are in group 15, and include N, P, As, Sb, and Bi.
(c) s^2p^6 is a *p* block configuration that corresponds to the general notation ns^2np^6. Elements having this configuration are in group 18, and include Ne, Ar, Kr, Xe, and Rn.

Check Your Solution

The link between *n* value and period, and the match between the valence electrons for each element and the general notation for elements in that group have been correctly applied.

13. Problem

For each of the elements below, use the aufbau principle to write the full and condensed electron configurations and draw partial orbital diagrams for the valence electrons of their atoms. You may consult the periodic table in Appendix C, or any other periodic table that omits electron configurations.

(a) potassium
(b) nickel
(c) lead

Solution

Element	Z	Electron configuration	Condensed configuration	Orbital diagram	Valence shell
(a) K	19	$1s^22s^22p^63s^23p^64s^1$	[Ar] $4s^1$	↑	$4s$
(b) Ni	28	$1s^22s^22p^63s^23p^64s^23d^8$	[Ar] $4s^23d^8$	↑↓	$4s$
(c) Pb	82	$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^2$ $4d^{10}5p^66s^24f^{14}5d^{10}6p^2$	[Xe]$6s^24f^{14}5d^{10}6p^2$	↑↓ ↑ ↑	$6s\,6p$

Check Your Solution

For each electron configuration, the aufbau principle has been followed. In the orbital diagram representing the valence electrons, Hund's Rule has been correctly followed.