

## Chapter 5 A Closer Look at Matter

### 5.1 Early Models of the Atom — Dalton to Rutherford

#### Warm Up, p. 234

1.

Subatomic Particle	Relative Charge	Relative Mass	Location in Atom
proton	1+	1 amu	inside nucleus
neutron	0	1 amu	inside nucleus
electron	1-	0	outside nucleus

2. The two subatomic particles responsible for virtually all of an atom's mass are the **proton** and the **neutron**.

3. Most of an atom's volume is actually **empty space**.

#### Quick Check, p. 235

1. Robert Boyle is credited with introducing the scientific method and with insisting on experimentation as a criterion to gather true knowledge about the world around us.

2. Greek philosophers such as Democritus relied largely on intuition whereas Dalton based his conclusions about the nature of matter on a large body of experimental data.

3. Dalton believed atoms to be the smallest bits of matter that existed. We now know that even smaller (subatomic) particles called protons, neutrons, and electrons are components of atoms.

#### Quick Check, p. 238

1. As the  $e/m$  ratio for the electron =  $1.76 \times 10^8 \text{ C/g}$  and the electron's charge as determined by Millikan is  $1.602 \times 10^{-19} \text{ C}$ , the following equation allows us to calculate electron mass:

$$\text{electron mass} = \frac{1.602 \times 10^{-19} \text{ C}}{1.76 \times 10^8 \text{ C/g}} = \mathbf{9.10 \times 10^{-28} \text{ g}}$$

2. (a) The faint glow coming from the cathode and the shadow of the anode appearing at the end of the discharge tube opposite to the cathode both suggested to Crookes that the mysterious ray originated at the cathode.

(b) The deflection of the cathode ray by a magnetic field suggested that the ray was actually a stream of charged particles.

(c) A beam of charged particles originating at the negative terminal and travelling toward the positive terminal is likely a beam of *negatively* charged particles.

3. (a) The atom is not the smallest particle of matter but is itself composed of even smaller subatomic particles.  
 (b) The subatomic components that make up the atom are oppositely charged entities.

**Quick Check, p. 241**

1. As only a tiny fraction of the alpha particles striking the atoms of the gold foil were deflected through large angles or even reflected back, Rutherford concluded that the part of the atom capable of deflecting those alpha particles, namely the nucleus, must be tiny compared to the actual size of the atom.
2. Rutherford's discovery of the nucleus occurred after the electron's negative charge and very small mass had been determined. He knew that a positive component of the atom must exist because an intact atom had no charge. He also knew that alpha particles were very massive and positively charged and so the only thing capable of significantly altering their course must also be massive and positive. It made sense therefore that the tiny nucleus should not only be where most of the atom's mass existed, but should also be positively charged.
3. In Rutherford's analogies, the reference to the  
 (a) "15-inch shell" represented the fast moving, massive alpha particles striking the atoms of the gold foil.  
 (b) "piece of tissue paper" represented what Rutherford originally considered the atoms of the gold foil to resemble according to Thomson's plum-pudding model of the atom.

4. The relative size of the atom compared to the nucleus is given by:

$$\frac{10^{-10} \text{ m}}{10^{-14} \text{ m}} = 10^4$$

Therefore, if the diameter of the nucleus was 2.0 cm, then the size of the atom would be given by:

$$2.0 \text{ cm} \times 10^4 = \mathbf{2.0 \times 10^4 \text{ cm}} \text{ (or } \mathbf{2.0 \times 10^2 \text{ m})}$$

This is about the length of two football fields!

**Practice Problems — Atomic Number, Mass Number, and Neutron Number, p. 243**

1. (a) 20 protons, 20 neutrons and 18 electrons.  
 (b) 52 protons, 76 neutrons, and 54 electrons.  
 (c) 24 protons, 28 neutrons, and 21 electrons.

2. (a)  ${}_{35}^{79}\text{Br}$       (b)  ${}_{79}^{197}\text{Au}$       (c)  ${}_{90}^{232}\text{Th}$

3.

Notation	Atomic Number	Mass Number	Number of Protons	Number of Neutrons	Number of Electrons
${}_{42}^{95}\text{Mo}$	42	95	42	53	42
${}_{34}^{79}\text{Se}^{2-}$	34	79	34	45	36
${}_{47}^{109}\text{Ag}^{+}$	47	109	47	62	46
${}_{90}^{232}\text{Th}^{4+}$	90	232	90	142	86
${}_{83}^{209}\text{Bi}$	83	209	83	126	83

**Practice Problems — Using Mass Spectrometry Data, p. 246**

1. Using the mass spectrometry data provided we can write:

$$(68.9256 \text{ amu})(0.6011) + (70.9247)(0.3989) = \mathbf{69.72 \text{ amu}}$$

The atomic mass of gallium is 69.72 amu.

2.

Most Deflection $\longrightarrow$ Least Deflection		
${}_{10}^{20}\text{Ne}^{1+}$	${}_{10}^{21}\text{Ne}^{1+}$	${}_{10}^{22}\text{Ne}^{1+}$

3. Let's ignore the units for now. The calculation is given by:

$$(x)(34.9689) + (1-x)(36.9659) = 35.453$$

$$34.9689x + 36.9659 - 36.9659x = 35.453$$

$$36.9659x - 34.9689x = 36.9659 - 35.453$$

$$x = 0.7576 \text{ and}$$

$$1 - x = 0.2424$$

Expressed as a percent, these become 75.76% and 24.24% respectively. Therefore the percent abundance of chlorine-35 is **75.76%** and the percent abundance of chlorine-37 is **24.24%**.

**5.1 Review Questions, p. 248**

1. (a) John Dalton viewed the atom as the smallest particle of matter resembling a solid sphere with no charged subatomic particles.

(b) Thomson's atom resembled a muffin with tiny negative electrons embedded in a positive amorphous "dough".

(c) Ernest Rutherford believed that an atom's volume was mostly empty space. A tiny dense positive nucleus at the centre contributed virtually all of the mass but none of the atom's volume and a cloud of orbiting electrons defined the atom's volume but contributed virtually none of the mass.

2. Thomson calculated a charge-to-mass ratio for the electron that was almost 2000 times greater than that of a positive hydrogen ion – the smallest particle known at the time. It made sense to Thomson that the magnitude of the two opposite charges should be equal which meant that the mass of the electron must be approximately 2000 less than the smallest atom.

3. The discovery of radioactivity led Rutherford to eventually discover alpha particles. Alpha particles were then used by Rutherford as a "probe" to test the internal structure of the atom. This experiment culminated in his discovery of the nucleus and his nuclear model of the atom.

4. Prior to the experiment, Rutherford believed in the "plum pudding" model of the atom proposed by J.J. Thomson. Nothing in Thomson's atom was capable of significantly altering the path of the massive fast-moving positive alpha particles, let alone turning them back towards their source.

5.

Notation	Name	Atomic Number	Mass Number	Number of Protons	Number of Neutrons	Number of Electrons
$^{131}_{54}\text{Xe}$	xenon-131	54	131	54	77	54
$^{210}_{84}\text{Po}$	polonium-210	84	210	84	126	84
$^{237}_{93}\text{Np}$	neptunium-237	93	237	93	144	93
$^{247}_{97}\text{Bk}$	berkelium-247	97	247	97	150	97
$^{257}_{100}\text{Fm}$	fermium-257	100	257	100	157	100

6.	(a) An atom with 78 protons and 117 neutrons	$^{195}_{78}\text{Pt}$	7.	(a) A 3+ cation with 80 electrons and 117 neutrons	$^{200}_{83}\text{Bi}^{3+}$
	(b) An atom with a mass of 237 having 90 electrons	$^{237}_{90}\text{Th}$		(b) A 1- anion with 54 electrons and 78 neutrons	$^{131}_{53}\text{I}^{1-}$
	(c) An atom with 69 electrons and 100 neutrons	$^{169}_{69}\text{Tm}$		(c) An ion with 66 electrons, 69 protons and 100 neutrons	$^{169}_{69}\text{Tm}^{3+}$

8. The calculation for the atomic mass of magnesium:

$$(23.9850 \text{ amu})(0.7899) + (24.9859 \text{ amu})(0.1000) + (25.9826 \text{ amu})(0.1101) = \mathbf{24.31 \text{ amu}}$$

9. Although the space surrounding the nucleus of atoms is not filled with matter, it is filled with negative electric fields produced by the electron cloud. This means that when atoms get close to each other, they exert repulsive forces that prevent matter from “collapsing” and give us the “illusion” of solid matter.

10. The calculation for the percent abundances of the two naturally occurring isotopes, ignoring the units:

$$\begin{aligned} (x)(106.9051) + (1-x)(108.9045) &= 107.87 \\ 106.9051x + 108.9045 - 108.9045x &= 107.87 \\ 1.9994x &= 1.0345 \\ x &= 0.5174 \text{ and} \\ 1-x &= 0.4826 \end{aligned}$$

Therefore, the percent abundance of silver-107 is **51.74%** and the percent abundance of silver-109 is **48.26%**.

## 5.2 Quantum Theory and the Bohr Model of the Atom

### Warm Up, p. 250

1. (a) Waves are disturbances that move through space and carry energy. They have a wavelength, a frequency, and an amplitude that can change by any amount, that is continuously.

(b) The two waves might pass through and interfere with each other. They might cancel each other out or result in a single larger wave.

2. (a) Solid objects are particles with definite boundaries whose amounts can only be in whole number quantities rather than any value.

(b) The two particles would collide with and bounce off each other and each would end up moving in a different direction with a different speed.

3. Students' answers will vary.

**Quick Check, p. 251**

1. For something to be considered “quantized” means it possesses particle-like properties. It exists in discrete whole number quantities and is considered to be “discontinuous”.
2. Examples of things considered to be quantized include eggs in a basket, students in a classroom, floors in an office building, and stairs in a staircase.
3. No. Only whole number packets of energy could ever be absorbed or emitted by an object.

**Quick Check, p. 255**

1. Hydrogen’s bright line spectrum consists of a series of four coloured lines (red, blue-green, blue, and violet) against a black background.
2. Excited electrons return to the second energy level ( $n = 2$ ) from higher levels by emitting specific amounts of energy in the visible spectrum. As electrons return from levels 6, 5, 4, and 3 down to level 2, a line is seen for each transition.
3. Any electron transitions from excited states down to the  $n = 1$  energy level generates an emission line in the ultraviolet region of the electromagnetic spectrum.

**Practice Problem, p. 256**

The equation  $\Delta E = b (1/n_1^2 - 1/n_2^2)$  is used to calculate the energy released. The value of the constant  $b = 2.18 \times 10^{-18}$  J and  $n_1 = 2$  and  $n_2 = 4$ . The calculation is:

$$\Delta E = 2.18 \times 10^{-18} \text{ J } (1/2^2 - 1/4^2) = \mathbf{4.08 \times 10^{-19} \text{ J}}$$

**5.2 Activity, p. 257****Results and Discussion**

1. As both of the visible lines in sodium’s emission spectrum are yellow, we would expect that the light emitted from a sodium vapour lamp would also appear yellow to the unaided eye.
  2. The letters: **l a s e r** stand for **l**ight **a**mplification by **s**timulated **e**mission of **r**adiation.
- Some applications of lasers in our modern society include:

*Medical*

1. “No-touch” surgery in removal of brain and spinal cord tumors
2. Eye surgery
3. Cosmetic surgery in removal of scars, sunspots, wrinkles, birthmarks, and tattoos
4. Tooth whitening and oral surgery

*Commercial and Industrial*

1. Guidance systems
2. Welding and cutting metals
3. Barcode readers
4. Engraving
5. Holography
6. Optical communications (over optical fibres or in free space)
7. Levels used in surveying and range finding

8. Laser printers
9. Storing and retrieving data in optical discs
10. Laser pointers
11. Writing subtitles onto motion picture films
12. Laser alignment
13. Laser lighting displays

*Scientific*

1. Atmospheric remote sensing
2. Laser-based LIght Detection and Ranging (LIDAR) used in geology, seismology, remote sensing, and atmospheric physics
3. Photochemistry
4. Laser cooling
5. Initiation of nuclear fusion
6. Microscopy

*Military*

1. Targeting and ranging
2. Defensive countermeasures
3. Communications and directed energy weapons
4. Laser sighting
5. Eye-targeted lasers

**5.2 Review Questions, p. 259**

1. The problem was that accelerating charges should radiate energy. As electrons orbiting the nucleus were accelerating charges, they should lose energy and thus instantly spiral into the nucleus resulting in atoms, and thus all matter, collapsing.
2. Planck's theory states that energy is quantized and can thus be emitted or absorbed only in whole numbers of discrete packets or bundles.
3. "Classical physics" viewed wave and particle behaviour as quite different and mutually exclusive from each other. The suggestion that a wave phenomenon such as energy could exhibit particle properties was difficult for most physicists to accept.
4. Albert Einstein's used Planck's theory to successfully explain the photoelectric effect. The phenomenon had baffled physicists up until then and could only be understood if light was considered to exist in discrete packets of energy.
5. Bohr reasoned that if energy was quantized and could thus exist in only certain amounts and not others, perhaps orbiting electrons in atoms could only possess certain amounts of energy and not others corresponding to orbits of only certain sizes.
6. According to Bohr, as long as an electron orbits the nucleus in an allowed orbit or stationary state, no energy is lost by that electron and so the atom does not collapse.
7. Because the amount of energy associated with each electron orbit is fixed, the difference in energy between those orbits is also fixed. Thus, when electrons "fall" from one excited state down to the  $n = 2$  energy level, the energy differences between  $n = 6$  and  $2$ ,  $n = 5$  and  $2$ ,  $n = 4$  and  $2$ , and  $n = 3$  and  $2$  are

always the same. These correspond only to four specific colours of light in the visible spectrum: violet, blue, blue-green, and red respectively and none others.

8. A continuous rainbow of colour would be observed with no dark regions.

9. Any energy emitted by electrons returning from excited states down to the  $n = 1$  ground state lowest energy level releases energy beyond the visible spectrum in the ultraviolet portion of the electromagnetic spectrum.

10. The calculation is given by:

$$\Delta E = 2.18 \times 10^{-18} \text{ J } (1/2^2 - 1/5^2) = 4.58 \times 10^{-19} \text{ J}$$

11. The calculation is given by:

$$1/\lambda = (1.097 \ 30 \times 10^7 \text{ m}^{-1}) (1/2^2 - 1/5^2)$$

$$1/\lambda = 2.3043 \times 10^6 \text{ m}^{-1}$$

$$\lambda = 433.97 \text{ nm}$$

This corresponds to the wavelength of blue light and thus represents the blue line in hydrogen's visible emission spectrum.

### 5.3 Beyond Bohr — The Quantum Mechanical Model of the Atom (Extension)

#### Warm Up, p. 261

1. Bohr's atomic theory suggested that the energies of orbiting electrons in atoms were quantized and so could have only certain values and not others. Bohr called these allowed orbits "stationary states."

2. The energies of electrons in atoms can only be certain amounts and not others, which correspond to electron orbits of only certain sizes and not others. When energy is absorbed by atoms, electrons are "excited" to higher energy levels corresponding to larger orbits further from the nucleus. No energy is lost when electrons exist in any of the allowed orbits. Energy is only emitted if excited electrons move from a higher energy orbit to a lower one and thus emit a specific amount of energy corresponding to the exact difference in energy between those two orbits or stationary states.

#### Quick Check, p. 264

1. Bohr was the first to suggest that the energies of electrons in atoms were quantized and explained how emission line spectra were produced.

2. Bohr's atomic model worked for hydrogen but was unable to predict the line spectra emitted by multi-electron atoms.

3. Particles large enough for us to see have wave natures far too insignificant for us to detect, let alone measure.

#### Practice Problems — Using de Broglie's Equation, p. 265

1. Use the equation:  $\lambda = h/mv$

The calculation is given by:

$$\lambda = \frac{6.626 \times 10^{-34} \text{ J s}}{(6.64 \times 10^{-27} \text{ kg})(1.6 \times 10^7 \text{ m/s})}$$



$$\lambda = 6.2 \times 10^{-15} \text{ m}$$

2. The above equation is manipulated to solve for  $m$ . This gives:  $m = h/\lambda v$

The calculation is given by:  $m = \frac{6.626 \times 10^{-34} \text{ kgm}^2/\text{s}}{(1.6 \times 10^7 \text{ m/s})(4.86 \times 10^{-7} \text{ m})}$  (note units for  $J \cdot s$ )

$$m = 8.5 \times 10^{-35} \text{ kg}$$

3. Use the equation:  $\lambda = h/mv$ . The calculation is given by:

$$\lambda = \frac{6.626 \times 10^{-34} \text{ J s}}{(0.056 \text{ kg})(66.7 \text{ m/s})}$$
 (note  $\text{km/h}$  changed to

$\text{m/s}$ )

$$\lambda = 1.8 \times 10^{-34} \text{ m}$$

### Quick Check, p. 268

1. Heisenberg considered the electron to be a pure particle and Schrödinger viewed the electron as a wave. Each developed mathematical equations to describe the electron accordingly.

2. The Heisenberg uncertainty principle forces us to accept that we can only describe electron behaviour using the mathematics of probability as prescribed by Schrödinger's wave equation. This has sometimes been compared to a game of chance such as rolling dice.

3. Bohr's "orbit" was a definite path followed by an electron particle as it circled the nucleus. An "orbital" is a region around the nucleus where an electron, viewed as a particle-wave, with a particular amount of energy is most likely to be found.

### 5.3 Activity, p. 269

#### Results and Discussion

1. The above scenarios show us that, when a measurement technique or instrument significantly affects what is being measured, the accuracy of the measurement is seriously compromised. This applies in quantum systems, that is, on the subatomic scale. Because the mass of the electron is so small, any attempt to measure its position or momentum (where it is or where it's going) so significantly affects those quantities that we can never specify both simultaneously.

### 5.3 Review Questions, p. 270

1. The Bohr model of the atom saw an electron as a pure particle. De Broglie suggested that the electron could also have wave properties.

2. By viewing the electron as having wave properties, de Broglie's hypothesis suggests that the only stable energy states for electrons in atoms are those that allow a standing electron wave to be maintained.

3. The wave properties of particles large enough to see are too insignificant to detect, let alone measure. This means that the wave nature of macroscopic particles can be ignored because their particle properties are virtually all that exist.

4. The wave nature of subatomic particles is so significant that it must be considered when describing their behaviour.

5. (a) As a particle's mass increases, its wave nature decreases.

(b) As a particle's mass decreases, its wave nature increases.

6. Planck insisted that energy, previously considered to be only a wave phenomenon, was quantized and thus possessed particle properties. De Broglie suggested that electrons, previously considered as being only particles, possessed wave properties.

7. The two theories, taken together, mean that perhaps pure waves and pure particles don't exist. Rather, particle-wave duality is a more accurate way of viewing the world.

**Note:** The original question 8 in the 1st edition has been omitted from the 2nd edition. The original questions 9 to 11 are now numbered 8 to 10, as below. The answer to question 8 in the 1st edition was: (a) By "He", Einstein was referring to the God of his belief. (b) Einstein was referring to a fundamental aspect of quantum mechanics that states that the laws of probability govern and limit our ability to describe matter at its most basic level.

8. There are two ways to answer this question. On a qualitative level, macroscopic objects such as cars and baseballs are too large for their position and motion to be significantly changed by what we use to measure those with. In other words, we *can* effectively state simultaneously both where macroscopic moving particles are located and where they're going.

Another way to answer this would be to note that because the magnitude of the constant  $h/2\pi$  in Heisenberg's Principle is so small, the magnitudes of the uncertainties in position and motion ( $\Delta x$  and  $m\Delta v$  respectively) are also very small compared to the sizes associated with macroscopic particles. We can, therefore, state both where "large" particles are and where they're going at the same time within our familiar frame of macroscopic reference.

9. De Broglie's equation is manipulated to solve for  $m$ . This gives:  $m = h/\lambda v$

The calculation is given by:

$$m = \frac{6.626 \times 10^{-34} \text{ kgm}^2/\text{s}}{(3.0 \times 10^6 \text{ m/s})(1.0 \times 10^{-12} \text{ m})} \quad (\text{note units for } J \text{ s})$$

$$m = 2.2 \times 10^{-28} \text{ kg}$$

10. (a) Niels Bohr was the first to establish the quantized energy states associated with electrons in atoms by applying Planck's quantum theory to atomic structure.

(b) Louis de Broglie was the first to suggest that particles could possess a wave nature which supplied a reason for the allowed energy states of electrons in the Bohr model of the atom.

(c) Erwin Schrodinger was the first to mathematically treat the electron as a wave and his wave equation allows us to determine the most probable locations for electrons of particular energies in atoms.

(d) Heisenberg's uncertainty principle forces us to accept that there is a limit to what can know for sure about matter at its most basic level, and compels us to accept the compromise of probability when describing electron behaviour.

## 5.4 Applying Quantum Mechanics to Electrons in Atoms, p. 271

### Warm Up, p. 271

1. Quantum mechanics retains the portion of Bohr's theory that quantized the energy associated with electrons in atoms.

2. Quantum mechanics defines an orbital as that region in three-dimensional space around a nucleus where an electron with a particular amount of energy is most likely to be found.

3. Heisenberg's uncertainty principle tells us that we cannot know simultaneously both where an electron is and where it's going. We are thus forced to employ Schrödinger's wave equation to identify for us the most probable location for electrons because we are unable to describe their behaviour with certainty.

### Quick Check, p. 275

1. The principle quantum number ( $n$ ) specifies the size of the atomic orbital.

The second or angular momentum quantum number ( $l$ ) tells us the shape of the atomic orbital.

The third or magnetic quantum number ( $m_l$ ) tells us the orientation in space of the atomic orbital.

2. (a) A 2s orbital is larger than a 1s orbital. This tells us that an electron in a 2s orbital is likely to spend most of its time further from the nucleus than an electron in a 1s orbital.

(b) Although a 2p<sub>x</sub> and a 2p<sub>y</sub> orbital have similar sizes and shapes, and energies, they are oriented at right angles or are *orthogonal* to each other.

3. A hydrogen electron in the fourth energy has 16 different orbitals to choose from — all at equal energy. They are: one 4s, three 4p, five 4d, and seven 4f orbitals. The notation  $n^2$  tells us how many orbitals are available in each energy level " $n$ ".

### Quick Check, p. 277

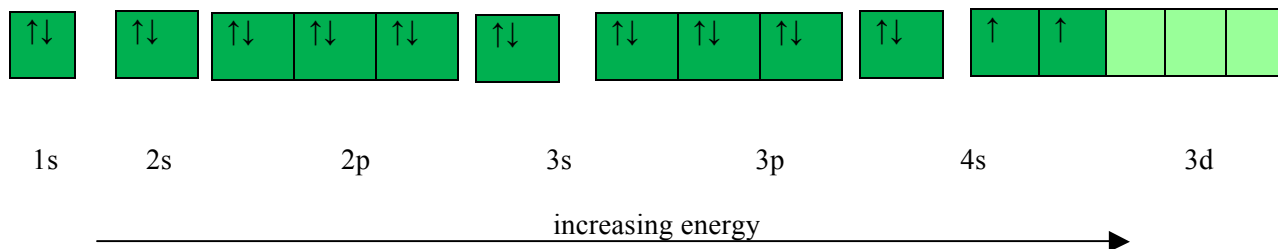
- The fourth or spin quantum number ( $m_s$ ) tells us the spin on an electron.
- If two electrons have the same *first three* quantum numbers, then those electrons must reside in the **same orbital**. If two electrons occupy the same orbital, then their spins, and thus their 4<sup>th</sup> quantum numbers, must be different.
- The maximum number of electrons that can exist in the  $n^{\text{th}}$  energy level is given by:  $2n^2$ . Therefore we can write the following table:

Energy Level ( $n$ )	Maximum Number of Electrons ( $2n^2$ )
1	2
2	8
3	18
4	32

### Practice Problem, p. 280

Full electron configuration for titanium:  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$

Orbital diagram for titanium:



### 5.4 Review Questions, p. 283

- Quantum mechanics tells us that a 1s orbital represents the (spherical) region in three-dimensional space around hydrogen's nucleus where its single electron is most likely located rather than a path that the electron follows.
- A 1s orbital is smaller than a 2s orbital. This tells us that an electron in a 2s orbital possessing energy equal to  $n = 2$  is more likely to be located further from the nucleus than an electron in a 1s orbital possessing energy equal to  $n = 1$ .
- The first difference is size. Any 3p orbital is larger than any 2p orbital. The "3" tells us that an electron in a 3p orbital possesses more energy than an electron in an orbital that begins with the number "2" such as a 2p orbital. This means that a 3p electron is more likely to be found further from the nucleus than a 2p electron. The other difference is associated with orientation in space. The lobes of a  $2p_x$  orbital are oriented at right angles to the lobes of a  $3p_y$  orbital.

4. No p orbital density at the nucleus means that there is no chance of finding the electron at that point in space. Another way of saying this is to say that the amplitude of the electron wave at the nucleus is zero. This is called a node.

5. The maximum number of electrons that can exist in the “ $n^{\text{th}}$ ” energy level is given by  $2n^2$ . Therefore, the first four energy levels can accommodate a maximum of: 2, 8, 18, and 32 electrons respectively.

6. Hund’s rule tells us that each of those three 3p electrons is in a different equal energy orbital: one in the  $3p_x$ , one in the  $3p_y$ , and one electron in the  $3p_z$ . Hund’s rule also tells us that each of those electrons is spinning in the same direction.

7. Use the periodic table to complete the following table:

Atom or Ion	Full Electron Configuration	Core Notation
<b>Ge</b>	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$	$[\text{Ar}] 4s^2 3d^{10} 4p^2$
<b>Zn<sup>2+</sup></b>	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$	$[\text{Ar}] 3d^{10}$
<b>Sr</b>	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2$	$[\text{Kr}] 5s^2$
<b>Br<sup>-</sup></b>	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$	$[\text{Ar}] 4s^2 3d^{10} 4p^6$
<b>Sn</b>	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^2$	$[\text{Kr}] 5s^2 4d^{10} 5p^2$
<b>In<sup>3+</sup></b>	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 4d^{10}$	$[\text{Kr}] 4d^{10}$

8. (a) Use the periodic table to identify the neutral atoms having the following electron configurations:

Electron Configuration	Element Name
$[\text{Ne}] 3s^2$	magnesium
$[\text{Ar}] 4s^2 3d^5$	manganese
$[\text{Kr}] 5s^2 4d^{10} 5p^3$	antimony
$[\text{Xe}] 6s^2 4f^7$	europium

(b) Each of the four elements in the above table is located in a different position of the periodic table corresponding to a different highest energy sublevel being filled in the elements.

The “s block” is located on the far left side of the periodic table; the “d block” is located in the centre of the table; the “p block” is located on the far right side of the table; the “f block” is located in the bottom region of the periodic table.

9. (a) Each ion possesses 10 electrons.

(b) Electron configuration of all species is given by:  $1s^2 2s^2 2p^6$

(c) Neon possesses this configuration and all of the ions achieve a stable valence octet by becoming isoelectronic with neon.

10. (a)

Alkali Metals	Core Notation	# Outer Electrons	Halogens	Core Notation	# Outer Electrons
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lithium	[He] 2s <sup>1</sup>	1	fluorine	[He] 2s <sup>2</sup> 2p <sup>5</sup>	7
sodium	[Ne] 3s <sup>1</sup>	1	chlorine	[Ne] 3s <sup>2</sup> 3p <sup>5</sup>	7
potassium	[Ar] 4s <sup>1</sup>	1	bromine	[Ar] 4s <sup>2</sup> 3d <sup>10</sup> 4p <sup>5</sup>	7
rubidium	[Kr] 5s <sup>1</sup>	1	iodine	[Kr] 5s <sup>2</sup> 4d <sup>10</sup> 5p <sup>5</sup>	7

(b) All the members of a chemical family demonstrate similar chemical behaviour because they all have the same number of outer or valence electrons.

(c) As we move down a chemical family, the atoms get larger.

