Multiple Choice Questions

1.
In a cathode ray tube

A. electrons pass from the anode to the cathode.
B. electrons pass from the cathode to the anode.
C. protons pass from the anode to the cathode.
D. protons pass from the cathode to the anode.

2. Which of the following scientists developed the nuclear model of the atom?
A. John Dalton
B. Robert Millikan
C. J. J. Thomson
D. Henry Moseley
E. Ernest Rutherford

3. Rutherford's experiment with alpha particle scattering by gold foil established that
A. protons are not evenly distributed throughout an atom.
B. electrons have a negative charge.
C. electrons have a positive charge.
D. atoms are made of protons, neutrons, and electrons.
E. protons are 1840 times heavier than electrons.

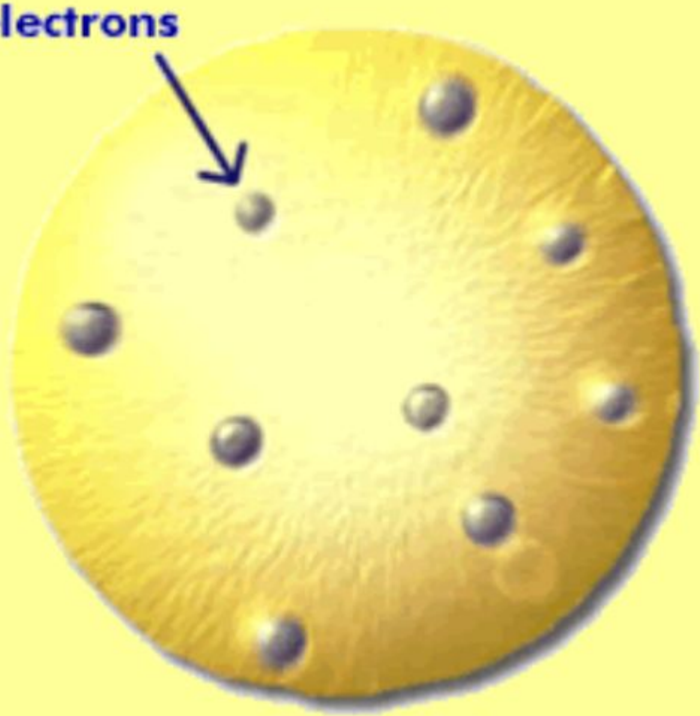
Sample Question

1. An atom consists of ________.
   A. protons only
   B. neutrons only
   C. protons and neutrons
   D. protons, neutrons and electrons
2. A neutron has ______ charge.
   A. positive
   B. negative
   C. zero
   D. variable
3. Among the following descriptions of element, atom and compound, which one is ***correct*** according to Dalton's Atomic Theory?
   A. Atoms cannot be further split into smaller portions.
   B. Compounds are composed of atoms of only one element.
   C. All atoms of a given element are identical, having the same size, mass and chemical properties.
   D. A chemical reaction can result in atoms creation or destruction.

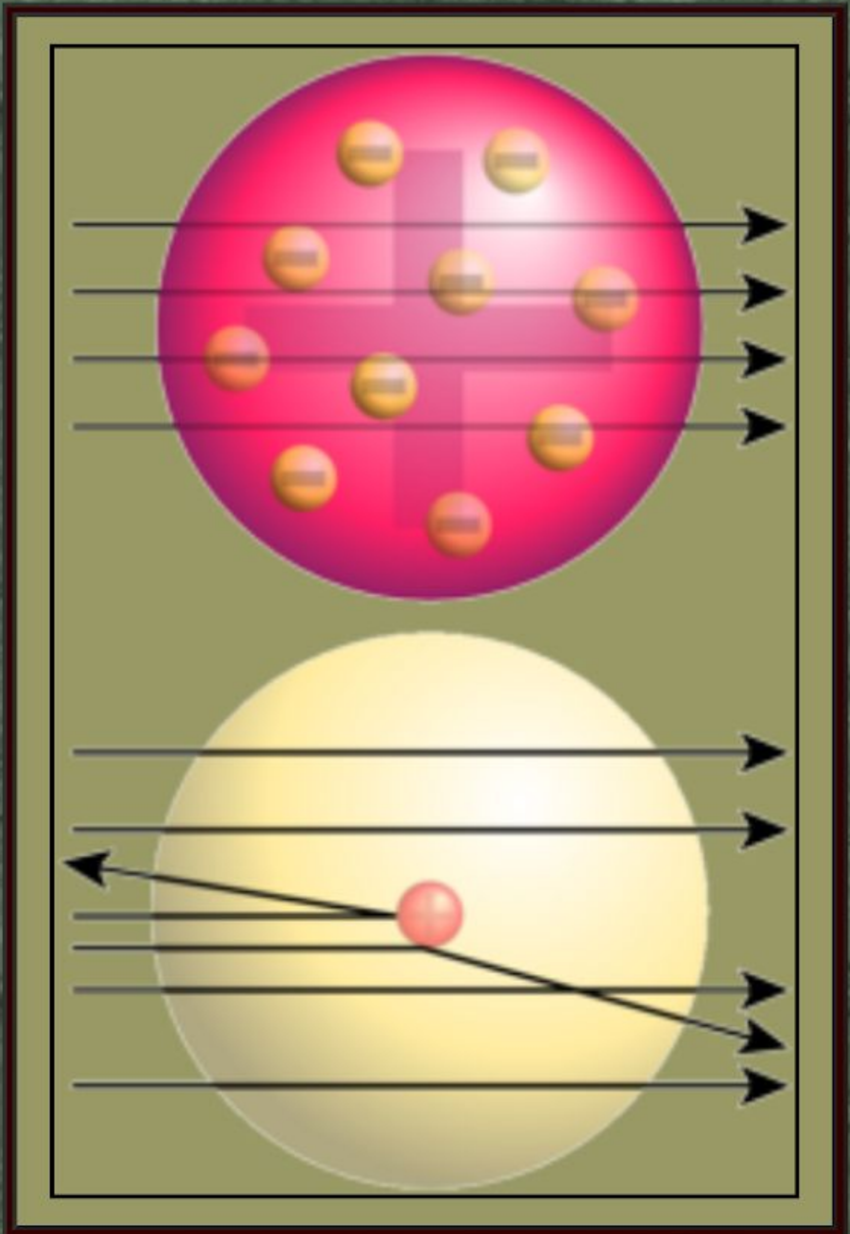
D. A chemical reaction can result in atoms creation or destruction.

4. Which of the following statements is ***wrong***?
   A. If two elements form more than a single compound, the masses of one element combined with a fixed mass of the second are in the ratio of small whole numbers.
   B. If two elements form more than a single compound, the masses of one element combined with a fixed mass of the second are in the ratio of any numbers.
   C. Atoms are neither created nor destroyed during physical or chemical processes.
   D. All samples of a compound have the same composition.
5. According to Rutherford's gold foil experiment, which of the following statement is ***wrong***?
   A. Undeflected straight-line paths exhibited by most of the particles.
   B. Some particles experienced slight deflections.
   C. Severe deflections of particles passing close to a nucleus.
   D. Reflections of particles from the foil is caused by approaching an electron head-on.
6. When very thin foils of gold are bombarded with α particles, which of the following phenomena can be observed?
   A. Few particles penetrated the foil undeflected.
   B. The majority of particles penetrated the foil undeflected.
   C. Most of particles suffered rather serious deflections as they penetrated the foil.
   D. The majority of particles did not pass through the foil at all, but bounced back in the direction from which they had come.

7. Most of the mass of the atom is due to ______; most of the volume of the atom is due to ______.
   A. nucleus; electrons
   B. neutrons; protons
   C. protons; nucleus
   D. electrons; nucleus
8. Isotopes have the same number of ______ but different numbers of ______.
   A. protons; neutrons
   B. neutrons; protons
   C. protons; electrons
   D. neutrons; electrons



JJ Thomson's **'Plum Pudding Model'** of the atom, a sphere of positive charge containing electrons



Practice Exercise How many protons, neutrons, and electrons are in the following isotope of copper: $^{63}_{29}Cu$?

- **Solution:**

- The atomic number is 29, so there are 29 protons. The mass number is 63, so the number of neutrons is 63-29=34. The number of electrons is the same as the number of protons, that is 29.

EXAMPLE 2.4 Interpreting Isotope Symbols

Determine the number of neutrons in each of the following isotopes.

(a) $^{238}_{92}U$ (b) ^{23}Na (c) hydrogen-3

Solution:

The number of neutrons is equal to the difference between the mass number (A) and the atomic number (Z). If the atomic number is not given in the symbol, look on the periodic table for the number that appears above the element symbol.

(a) $238 - 92 = 146$ neutrons
(b) $23 - 11 = 12$ neutrons
(c) $3 - 1 = 2$ neutrons

Which of the following are isotopes?

A. ^{14}C and ^{13}C

B. ^{14}C and ^{14}N

C. ^{14}N and $^{14}N^{3-}$

D. ^{12}C and ^{12}CO

E. ^{14}N and $^{14}N_2$

Complete the following chart, in order from left to right

Isotope	Mass Number	Protons	Neutrons	Electrons
^{14}N				

A. 14, 7, 7, 7
B. 14, 7, 14, 7
C. 7, 7, 7, 7
D. 7, 14, 7, 7
E. Some other answer

Isotope	Mass Number	Protons	Neutrons	Electrons
	40	19		19

A. ^{40}Zr, 21
B. ^{19}K, 40
C. ^{21}K, 19
D. ^{40}K, 21
E. ^{38}Sr, 19

Isotope	Mass Number	Protons	Neutrons	Electrons
		40	57	40

A. ^{97}Zr, 97
B. ^{40}Zr, 57
C. ^{57}La, 40
D. ^{97}Bk, 80
E. ^{80}Hg, 97

Solution:

(a) The atomic number (the number of protons) is 35, indicating the element symbol is Br. The mass number is the sum of the protons and neutrons, which is 79. The charge is 1− because there is one more electron than protons. The symbol is ${}^{79}_{35}\mathrm{Br}^-$.

(b) The atomic number is 13, so the element symbol is Al. The mass number is 27 (13 + 14). The charge is 3+ because there are three fewer electrons than protons. The symbol is ${}^{27}_{13}\mathrm{Al}^{3+}$.

(c) The atomic number is 47, so the element is silver (Ag). The mass number is 109 (47 + 62). The charge is 1+ because the number of electrons is one less than the number of protons. The symbol is ${}^{109}_{47}\mathrm{Ag}^+$.

For your reference

How can we describe the mass of an atom of an element? Although single atoms cannot be weighed on a balance, modern techniques such as *mass spectrometry* (Figure 2.15) can be used to determine individual atomic masses accurately. For example, the mass of a single hydrogen-1 atom is 1.67380×10^{-24} g. The mass of a carbon-12 atom is 1.99272×10^{-23} g, about 12 times the mass of a hydrogen-1 atom. Numbers as small as these are difficult to remember and use. It is more convenient to think of a carbon-12 atom as being about 12 times the mass of a hydrogen-1 atom. For this reason, scientists have devised a method for expressing masses of atoms in a more convenient way. They use the atomic mass unit (amu). This mass scale uses carbon-12 (^{12}C), the most abundant isotope of carbon, as the standard to which all other atoms are compared. Carbon-12 is assigned an atomic mass of exactly 12 atomic mass units, or 12 amu. One **atomic mass unit (amu)** is equal to one-twelfth *the mass of a carbon-12 atom:*

$$1 \text{ amu} = \tfrac{1}{12} \times \text{mass of 1 C-12 atom} = 1.6606 \times 10^{-24} \text{ g}$$

Answer

2) 70%

Solution

$$62.9X + 64.9\,(1-X) = 63.5$$

$$-2.0\,X = -1.4$$

$$X = 70\%$$

Solution:

The relative atomic mass of chlorine is 35.45 amu, as shown on the periodic table. Since the relative atomic mass is closer to the mass of ^{35}Cl (34.9689 amu) than the mass of ^{37}Cl (36.9659 amu), the ^{35}Cl isotope must be most abundant. (The actual percentages are 75.77% ^{35}Cl and 24.23% ^{37}Cl.)

As you are aware, "mass" is a physical property of matter. The mass of an atom is known as it's **Atomic Mass**. Atomic mass is calculated by adding the number of protons and neutrons in an atom. *The electrons are not counted because of their small size*. The unified atomic mass unit (symbol: u) or **Dalton** (symbol: Da) is the standard unit that is used for indicating mass on an atomic or molecular scale. The atomic mass of an atom will change depending on its amount of protons and neutrons and will be different for each individual element.

Atomic mass is also known as the **Atomic Weight,** with the abbreviation at.wt. Actually the two terms are not exact equivalent. "Weight" implies a force exerted in a gravitational field, which would be measured in units of force, like Newtons. The term "atomic weight" is in use for two centuries or more, so it continues to be used; however, to avoid any confusion, atomic weight is more commonly known now, as **relative atomic mass.** As the name indicates, it is a ratio of two masses. The mass of an atom is compared with that of an atom of Carbon-12. The relative atomic mass of Carbon-12 is taken to be 12. Relative masses have no units.

Mass number is simply a count of the total number of protons and neutrons in an atom's nucleus. The Mass Number of an atom is represented by the symbol A and is also known as the Nucleon Number.

What is optoelectronics?
The study and application of electronic devices that source, detect and control light

LEDs

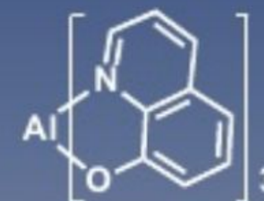
CRT

lasers

solar cells

PM

- the classical devices use inorganic materials: Si, GaN, Y_2O_2S:Eu, YAG:Nd
- 1987: Tang and van Slyke demonstrate the first organic optoelectronic device
- nowadays:



Properties of lanthanide ions

La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu

Ln^{III} ground state
$[Xe]4f^n$, n = 0..14

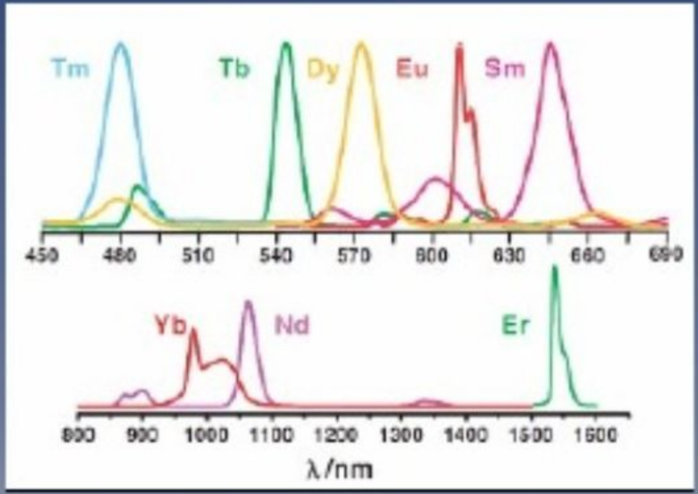
Shielding of 4f orbitals ⇨

- similar chemical properties
- electrostatic bonding
- variable geometry and CNs
- hard acid behaviour

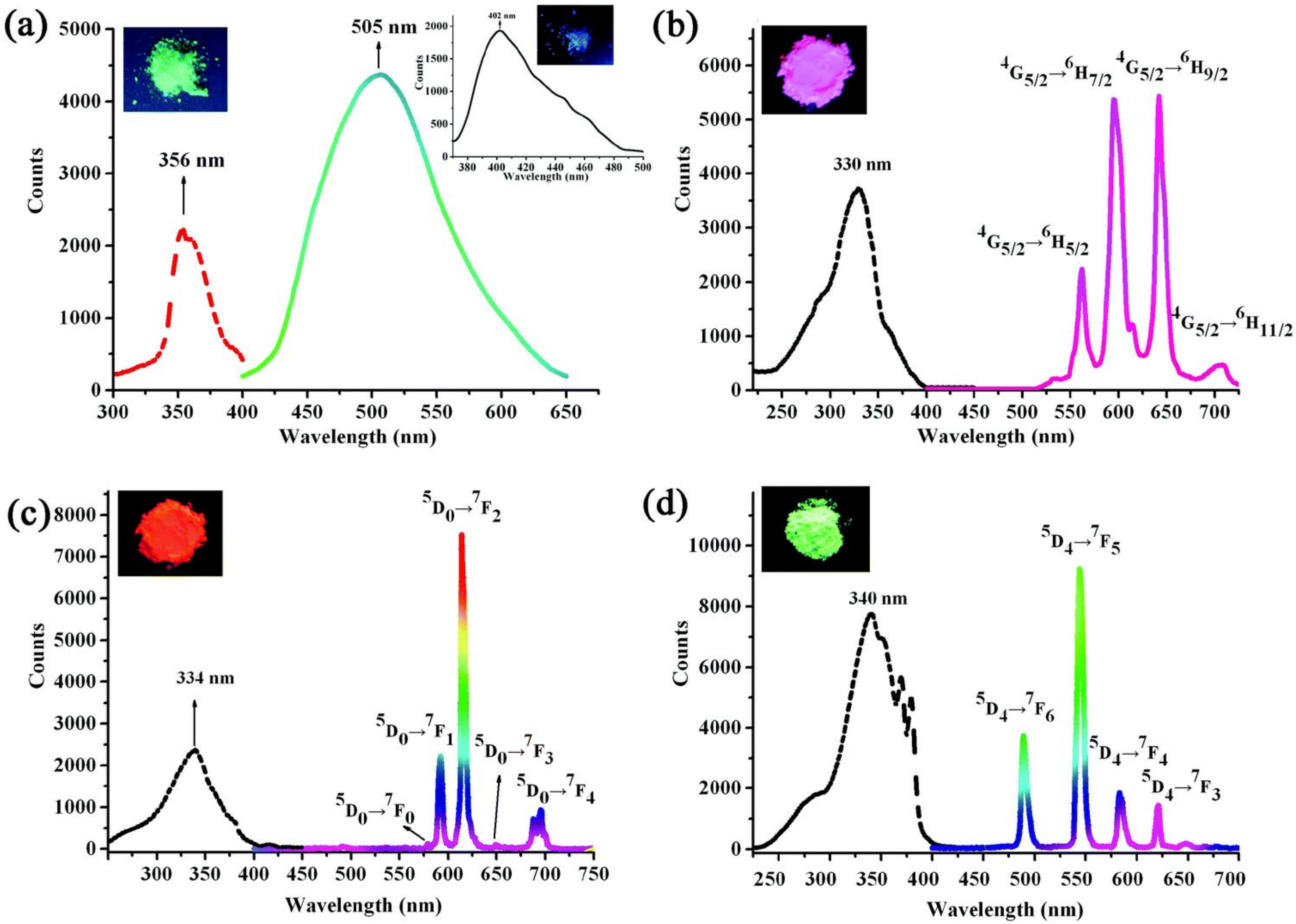
Fascinating optical properties:

- luminescence from f-f transitions
- characteristic emission for each ion
- narrow emission bands
- long excited-states lifetimes

blue → NIR

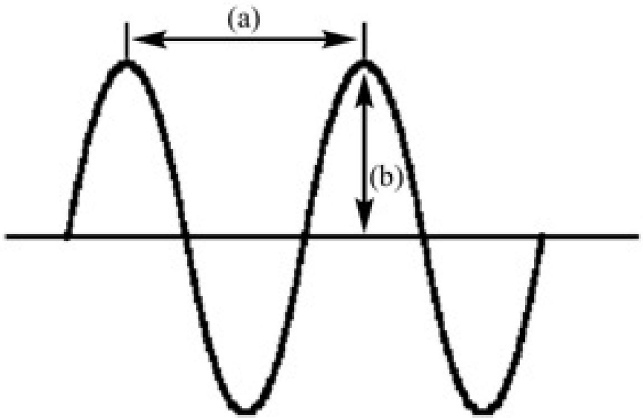


⇨ Applications in optoelectronics and bio-medicine



(a)
505 nm
402 nm
356 nm
Counts
Wavelength (nm)
(b)
330 nm
4G5/2→6H5/2
4G5/2→6H7/2
4G5/2→6H9/2
4G5/2→6H11/2
Counts
Wavelength (nm)
(c)
334 nm
5D0→7F0
5D0→7F1
5D0→7F2
5D0→7F3
5D0→7F4
Counts
Wavelength (nm)
(d)
340 nm
5D4→7F6
5D4→7F5
5D4→7F4
5D4→7F3
Counts
Wavelength (nm)

In the following diagram of a wave



A. (a) is amplitude and (b) is wavelength
B. (a) is frequency and (b) is amplitude
C. (a) is wavelength and (b) is frequency
D. (a) is amplitude and (b) is frequency
E. (a) is wavelength and (b) is amplitude

- This relationship is described by the universal wave equation

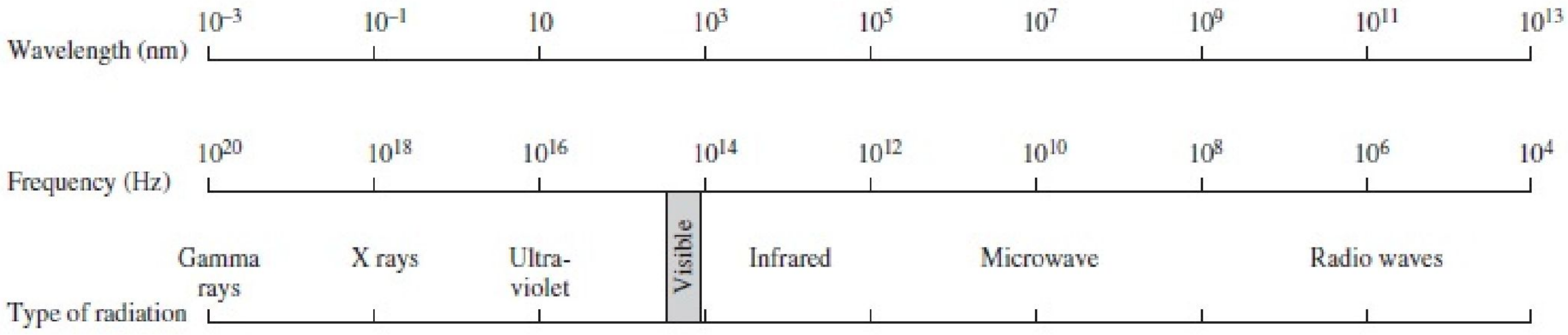
$$v = f\lambda$$

v – speed (m/s)

f – frequency (Hz or 1/s)

λ – wavelength (m)

Using the figure below, categorize electromagnetic radiation with a wavelength of 1.0×10^{-3} m.



A. Gamma rays
B. X rays
C. Ultraviolet
D. Infrared
E. Microwave

- **Solution:**
- Replacing *u* with c (the speed of light) gives
- $c = \lambda \times \nu$
- $\lambda = c/\nu$
- $= 3.00 \times 10^8$ m/s / 3.64×10^7 Hz
- $= 3.00 \times 10^8$ m/s / 3.64×10^7 /s
- $= 8.24$ m

Which of the following wavelengths of electromagnetic radiation has the highest energy?

A. 450. nm

B. 225 nm

C. 3.50×10^{-9} m

D. 8.40×10^{-7} m

E. 2.50×10^{-5} m

Solution:

Recall the relationship between wavelength, λ, and frequency, ν:

$$\boxed{\nu\lambda = c}$$

where c is the speed of light, 3.00×10^8 m/s. We know wavelength, so we can rearrange the equation to solve for frequency:

$$\nu = \frac{c}{\lambda}$$

Wavelength is given in units of nanometers, so we must convert to meters to cancel units with the units of speed of light (m/s). (For a review of unit conversions, see Math Toolbox 1.3.) The relationship between nanometers and meters is

$$1 \text{ m} = 10^9 \text{ nm}$$

This relationship can be written in the following two ways:

$$\frac{10^9 \text{ nm}}{1 \text{ m}} \quad \text{and} \quad \frac{1 \text{ m}}{10^9 \text{ nm}}$$

When we convert from nanometers to meters, we multiply the known value in nanometers by the conversion factor that has the nanometer units in the denominator to cancel units properly:

$$805 \cancel{\text{ nm}} \times \frac{1 \text{ m}}{10^9 \cancel{\text{ nm}}} = 8.05 \times 10^{-7} \text{ m}$$

$$\lambda = 8.05 \times 10^{-7} \text{ m}$$

Now that we have the appropriate units for wavelength, we can divide the speed of light by its value to get frequency:

$$\nu = \frac{c}{\lambda}$$

$$= \frac{3.00 \times 10^{8}\ \cancel{m}/s}{8.05 \times 10^{-7}\ \cancel{m}} = 3.73 \times 10^{14}/s \quad \text{or} \quad \boxed{3.73 \times 10^{14}\ \text{Hz}}$$

The energy of a photon can be calculated from frequency using the following equation:

$$\boxed{E_{\text{photon}} = h\nu}$$

where h is Planck's constant, 6.626×10^{-34} J•s. Since we have just calculated frequency, we can substitute values of h and ν into the equation:

$$E_{\text{photon}} = 6.626 \times 10^{-34}\ \text{J}\cdot\cancel{s} \times 3.73 \times 10^{14}\ \frac{1}{\cancel{s}} = \boxed{2.47 \times 10^{-19}\ \text{J}}$$

Units of seconds cancel, giving units of joules, which is a unit of energy. The energy of a single photon of this light is 2.47×10^{-19} J.

Practice Exercise The energy of a photon is 5.87×10^{-20} J. What is its wavelength (in nanometers)?

- **Solution:**

  $E = h \times c/ \lambda$

- $\lambda = h \times c/ E$
- $= 6.63 \times 10^{-34}\ (J \cdot s) \times 3.00 \times 10^{8}\ (m/s) / 5.87 \times 10^{-20}\ J$
- $= 3.39 \times 10^{-6}\ m$
- $= 3.39 \times 10^{3}\ nm$

$$\Delta E = E_f - E_i$$

$$= -R_H \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$$

$$= R_H \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right)$$

- Solution:

$$\Delta E = R_H\left(\frac{1}{n_i 2} - \frac{1}{n_f 2}\right)$$
$$= 2.18\text{x}10^{-18}J\left(\frac{1}{6^2} - \frac{1}{4^2}\right)$$
$$= -7.57\text{x}10^{-20}J$$

- The negative sign indicates that this is energy associated with an emission process. To calculate the wavelength, we will omit the minus sign because the wavelength of photon must be positive.

$$\lambda = \frac{c}{v} = \frac{ch}{\Delta E}$$
$$= \frac{(3.00 \times 10^8 m/s)(6.63 \times 10^{-34} J \cdot s)}{7.57 \times 10^{-20} J}$$
$$= 2.63 \times 10^{-6} m = 2.63 \times 10^3 nm$$

Solution:

(a) In the hydrogen line spectrum in Figure 7.10, the wavelength of red light is given as 656 nm. Knowing the wavelength, λ, we can find E_{photon} using the equation that relates wavelength to photon energy:

$$E_{\text{photon}} = \frac{hc}{\lambda}$$

We must convert the wavelength units from nanometers to meters so units cancel, just as we did in Example 7.2:

$$656\ \cancel{\text{nm}} \times \frac{1\ \text{m}}{10^9\ \cancel{\text{nm}}} = 6.56 \times 10^{-7}\ \text{m}$$

Now we substitute Planck's constant, the speed of light, and the wavelength into the equation for photon energy:

$$E_{\text{photon}} = \frac{6.626 \times 10^{-34}\ \text{J} \cdot \cancel{\text{s}} \times 3.00 \times 10^8\ \frac{\cancel{\text{m}}}{\cancel{\text{s}}}}{6.56 \times 10^{-7}\ \cancel{\text{m}}} = 3.03 \times 10^{-19}\ \text{J}$$

(b) From the equation that relates photon energy to wavelength, we see that photon energy is *inversely* proportional to wavelength—the shorter the wavelength, the higher the photon energy. In Figure 7.10 we see that the violet line in the hydrogen spectrum has the shortest wavelength, so it has the highest photon energy. The highest photon energy, E_{photon}, corresponds to the largest electron transition ΔE, which is the difference in energy between the orbits. The $n = 6$ orbit is further in energy from the $n = 2$ orbit than are the other orbits, so the transition from $n = 6$ to $n = 2$ emits light of highest energy.

Extension to Higher Z

- The Bohr model can be extended to any ***single electron system***….must keep track of Z (atomic number).

$$E = -2.18x10^{-18}J\left(\frac{Z^2}{n^2}\right)$$

Z = atomic number

n = integer (1, 2, ….)

- Examples: He^{+} (Z = 2), Li^{+2} (Z = 3), etc.

Extension to Higher Z (cont.)

- **Example 2: At what wavelength will emission from n = 4 to n = 1 for the He$^+$ atom be observed?**

$$\Delta E = 2.18x10^{-18} J \left(\cancel{Z}^2\right)_{2}\left(\frac{1}{\cancel{n}^2_{initial}} - \frac{1}{\cancel{n}^2_{final}}\right)$$

$$\quad\quad 4 \quad\quad\quad 1$$

$$\Delta E = 2.18x10^{-18} J (4)\left(\frac{1}{16} - 1\right) = -8.16x10^{-18} J$$

$$\Delta E = 8.16x10^{-18} J = \frac{hc}{\lambda} \Rightarrow \lambda = 2.43x10^{-8} m = 24.3 nm$$

$$\lambda_H > \lambda_{He^+}$$

The electron in a hydrogen atom falls from an excited energy level to the ground state in two steps, causing the emission of photons with wavelengths of 2624 nm and 97.2 nm, respectively. What is the principal quantum number or shell of the initial excited energy level from which the electron falls?

A. 2
B. 3
C. 4
D. 6
E. 8

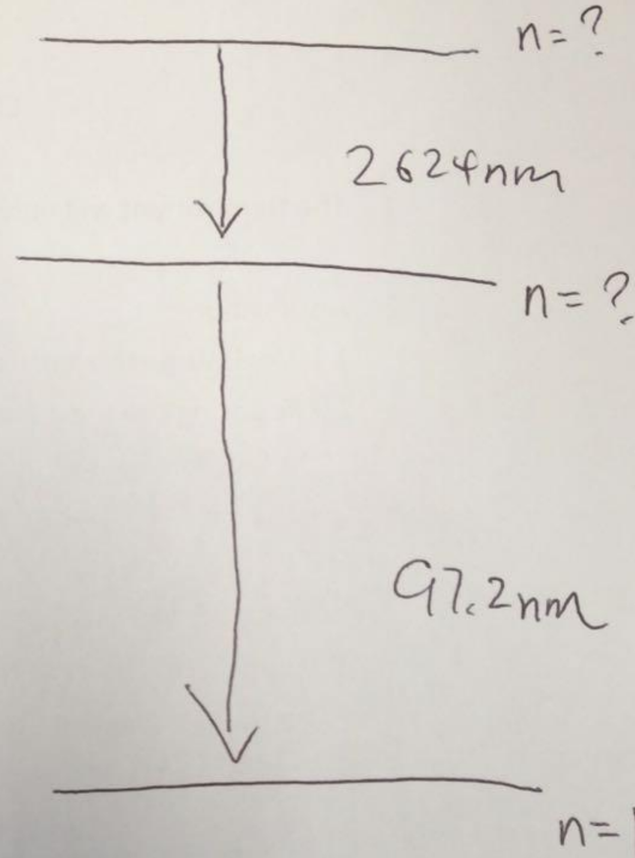
Bloom's Level: 4. Analyze
Difficulty: ***Difficult***
Gradable: automatic
Section: 07.03
Subtopic: Atomic Spectra (Bohr Model of the Atom)
Subtopic: Electromagnetic Radiation
Subtopic: Quantum Numbers
Topic: Quantum Theory and Atomic Structure

ΔE for 2624 nm

$$\Delta E = \frac{ch}{\lambda}$$

$$= \frac{3.00 \times 10^{8} \times 6.63 \times 10^{-34}}{2624 \times 10^{-9}}$$

$$= 7.58 \times 10^{-20}\ J$$



ΔE for 97.2 nm

$$\Delta E = \frac{ch}{\lambda}$$

$$= 2.046 \times 10^{-18}\ J$$

$$\Delta E = R_H \left(\frac{1}{n^2_i} - \frac{1}{n^2 f} \right)$$

for 97.2nm

$$-2.046 \times 10^{-18} = 2.18 \times 10^{-18} \left(\frac{1}{n^2_i} - \frac{1}{1^2} \right)$$

$$n_i = 4$$

$$\Delta E = R_H \left(\frac{1}{n^2_i} - \frac{1}{n^2 f} \right)$$

for 2624nm

$$-7.58 \times 10^{-20} = 2.18 \times 10^{-18} \left(\frac{1}{n^2_i} - \frac{1}{4^2} \right)$$

$$n_i = 6$$

The electron in a hydrogen atom falls from an excited energy level to the ground state in two steps, causing the emission of photons with wavelengths of 2624 nm and 97.2 nm, respectively. What is the principal quantum number or shell of the initial excited energy level from which the electron falls?

A. 2

B. 3

C. 4

D. 6

E. 8

Bloom's Level: 4. Analyze

*Difficulty: **Difficult***

Gradable: automatic

Section: 07.03

Subtopic: Atomic Spectra (Bohr Model of the Atom)

Subtopic: Electromagnetic Radiation

Subtopic: Quantum Numbers

Topic: Quantum Theory and Atomic Structure

$$\text{Moles} = \frac{\text{Mass}}{\text{Molar Mass}}$$

What is the difference between **molecular weight** and **molar mass**?

- In practice, nothing. **The units are different**:
  * molecular weight is in atomic mass units, because it's the mass of one molecule.
  * molar mass is in grams per mole, because it's the mass of a mole of molecules.

  But the numbers are the same.

Solution:

Consider a 100-g sample of malachite, so the masses of each element are numerically the same as the percent composition. Then calculate the moles of each element in 100 g of the compound. We use the molar masses to create equivalence expressions and convert them to ratios for the conversions. As in Example 4.9, the correct ratio has moles divided by grams:

$$1 \text{ mol Cu} = 63.55 \text{ g Cu} \qquad \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}}$$

$$1 \text{ mol C} = 12.01 \text{ g C} \qquad \frac{1 \text{ mol C}}{12.01 \text{ g C}}$$

$$1 \text{ mol H} = 1.008 \text{ g H} \qquad \frac{1 \text{ mol H}}{1.008 \text{ g H}}$$

$$1 \text{ mol O} = 16.00 \text{ g O} \qquad \frac{1 \text{ mol O}}{16.00 \text{ g O}}$$

We multiply the mass of each element by the corresponding conversion ratio to get moles of this element:

$$\text{mol Cu} = 57.48\ \cancel{\text{g Cu}} \times \frac{1\ \text{mol Cu}}{63.55\ \cancel{\text{g Cu}}} = 0.9045\ \text{mol Cu}$$

$$\text{mol C} = 5.43\ \cancel{\text{g C}} \times \frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}} = 0.452\ \text{mol C}$$

$$\text{mol H} = 0.91\ \cancel{\text{g H}} \times \frac{1\ \text{mol H}}{1.008\ \cancel{\text{g H}}} = 0.90\ \text{mol H}$$

$$\text{mol O} = 36.18\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 2.261\ \text{mol O}$$

To calculate the mole ratios, first identify the element present in the smallest amount in moles. In this example, it's carbon. Then divide the moles of each of the other elements by the moles of carbon:

$$\frac{\text{mol Cu}}{\text{mol C}} = \frac{0.9045 \text{ mol Cu}}{0.452 \text{ mol C}} = \frac{2.00 \text{ Cu}}{1.00 \text{ C}}$$

$$\frac{\text{mol H}}{\text{mol C}} = \frac{0.90 \text{ mol H}}{0.452 \text{ mol C}} = \frac{1.99 \text{ H}}{1.00 \text{ C}}$$

$$\frac{\text{mol O}}{\text{mol C}} = \frac{2.261 \text{ mol O}}{0.452 \text{ mol C}} = \frac{5.00 \text{ O}}{1.00 \text{ C}}$$

When the mole ratios are rounded to the nearest whole number, we see that one formula unit has one carbon atom, two copper atoms, two hydrogen atoms, and five oxygen atoms. From this information, we can write the empirical formula $Cu_2CO_5H_2$. [The formula is usually written as $Cu_2CO_3(OH)_2$, to show which ions are found in the mineral. Malachite contains two anions, hydroxide, OH^-, and carbonate, CO_3^{2-}. However, this cannot be deduced from the empirical formula for this compound.]

Example 2.4

DCM, Worked Examples

Write the formula of magnesium nitride, containing the Mg^{2+} and N^{3-} ions.

Strategy Our guide for writing formulas for ionic compounds is electrical neutrality; that is, the total charge on the cation(s) must be equal to the total charge on the anion(s). Because the charges on the Mg^{2+} and N^{3-} ions are not equal, we know the formula cannot be MgN. Instead, we write the formula as Mg_xN_y, where x and y are subscripts to be determined.

Solution To satisfy electrical neutrality, the following relationship must hold:

$$(+2)x + (-3)y = 0$$

Solving, we obtain $x/y = 3/2$. Setting $x = 3$ and $y = 2$, we write

$$Mg^{2+} \qquad N^{3-}$$

$$Mg_3N_2$$

Check The subscripts are reduced to the smallest whole number ratio of the atoms because the chemical formula of an ionic compound is usually its empirical formula.

Solution:

In both cases the charge on tin is not apparent from its position in the periodic table, so we'll have to deduce its charge in each compound from the anions.

(a) Based on its position in the periodic table, we predict that chloride ion has a 1− charge. For the compound to be electrically neutral, the tin must have a 2+ charge to counter the negative charges from the two chloride ions. We can now write the name for the compound by beginning with the metal name followed by the Roman numeral to represent its charge. Then we write the root of the nonmetal name for the anion with an *–ide* ending. The name of the compound is *tin(II) chloride.*

(b) Based on its position in the periodic table, we predict that oxide ion has a 2− charge. For the compound to be electrically neutral, the tin must have a 4+ charge to counter the four negative charges from the two oxide ions. We can now write the name for the compound by beginning with the metal name followed by the Roman numeral to represent its charge. Then we write the root of the nonmetal name for the anion with an *–ide* ending. The name of the compound is *tin(IV) oxide.*

Example 2.5

DCM, Worked Examples

Name the following compounds: (a) $Cu(NO_3)_2$, (b) KH_2PO_4, and (c) NH_4ClO_3.

Strategy Note that the compounds in (a) and (b) contain both metal and nonmetal atoms, so we expect them to be ionic compounds. There are no metal atoms in (c) but there is an ammonium group, which bears a positive charge. So NH_4ClO_3 is also an ionic compound. Our reference for the names of cations and anions is Table 2.3. Keep in mind that if a metal atom can form cations of different charges (see Figure 2.11), we need to use the Stock system.

Solution

(a) The nitrate ion (NO_3^-) bears one negative charge, so the copper ion must have two positive charges. Because copper forms both Cu^+ and Cu^{2+} ions, we need to use the Stock system and call the compound copper(II) nitrate.

(b) The cation is K^+ and the anion is $H_2PO_4^-$ (dihydrogen phosphate). Because potassium only forms one type of ion (K^+), there is no need to use potassium(I) in the name. The compound is potassium dihydrogen phosphate.

(c) The cation is NH_4^+ (ammonium ion) and the anion is ClO_3^-. The compound is ammonium chlorate.

Solution:

(a) The name tells us that iron has a 3+ charge in this compound, Fe^{3+}. Sulfide is the name of the monatomic ion of sulfur that, from its position in the periodic table, we expect to have a charge of 2−, S^{2-}. For the sum of the positive and negative charges to be equal, there must be three sulfide ions for every two iron(III) ions, giving a formula of Fe_2S_3.

(b) We know from the name that cobalt has a 2+ charge in this compound, Co^{2+}. Nitrate is the name of a polyatomic ion with a charge of 1− (NO_3^-). For the sum of the positive charges and the sum of the negative charges to be equal, there must be two nitrate ions for every one cobalt(II) ion, giving a formula of $Co(NO_3)_2$.

Example 2.8

DCM, Worked Examples

Write chemical formulas for the following molecular compounds: (a) carbon disulfide and (b) disilicon hexabromide.

Strategy Here we need to convert prefixes to numbers of atoms (see Table 2.4). Because there is no prefix for carbon in (a), it means that there is only one carbon atom present.

Solution (a) Because there are two sulfur atoms and one carbon atom present, the formula is CS_2.

(b) There are two silicon atoms and six bromine atoms present, so the formula is Si_2Br_6.

EXERCISE 9.1

Name these coordination complexes:

a. $Cr(NH_3)_3Cl_3$

b. $Pt(en)Cl_2$

c. $[Pt(ox)_2]^{2-}$

d. $[Cr(H_2O)_5Br]^{2+}$

e. $[Cu(NH_2CH_2CH_2NH_2)Cl_4]^{2-}$

f. $[Fe(OH)_4]^{-}$

Answer:

a. Triamminetrichlorochromium(III)

b. Dichloroethylenediamineplatinum(II)

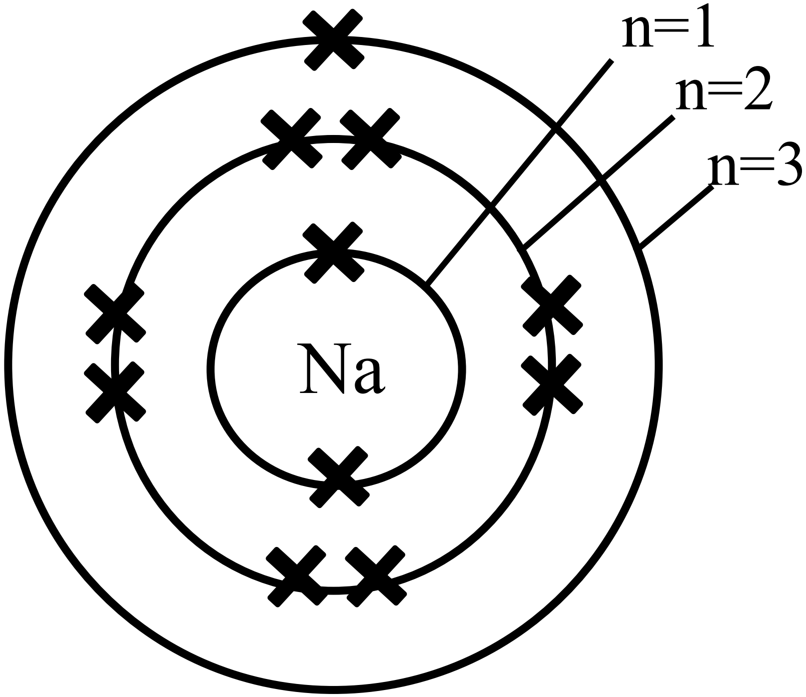
c. Bis(oxalato)platinate(II) or bis(oxalato)platinate(2−)

d. Pentaaquabromochromium(III) or pentaaquabromochromium(2+)

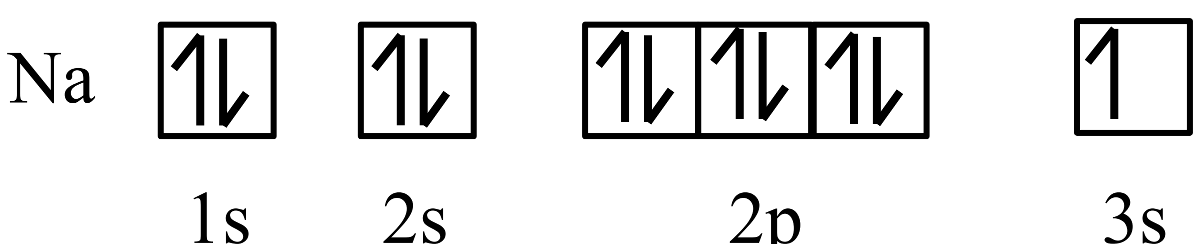
e. Tetrachloroethylenediaminecuprate(II) or tetrachloroethylenediaminecuprate(2−)

f. Tetrahydroxoferrate(III) or tetrahydroxoferrate(1−)

Simplified Model



Orbital diagram



- **Solution:**
- The number given in the designation of the subshell is the principle quantum number, so in this case **n = 3**. The letter designates the type of orbital. Because we are dealing with **p orbitals**, ***l*** **=1**. The values of m_l can vary from ***–l* to *l*.** Therefore, m_l can be **-1, 0, 1**.

- **Check:**
- The values of *n* and *l* are fixed for 3p, but m_l can have any one of the three values, which correspond to the three p orbitals.

- **Solution:**
- The principal quantum number n is 5 and the angular momentum quantum number l must be 1 (because we are dealing with a p orbital).
- For l=1, there are three values of m_l given by -1, 0, and 1. Because the electron spin quantum number m_s can be either +1/2 or -1/2, we conclude that there are six possible ways to designate the electron using the (n, l, m_l, m_s) notation:

(5, 1, -1, +1/2)	(5, 1, -1, -1/2)
(5, 1, 0, +1/2)	(5, 1, 0, -1/2)
(5, 1, 1, +1/2)	(5, 1, 1, -1/2)

Solution:

(a) The orbital diagram is correct for beryllium.

(b) The orbital diagram for oxygen is not correct. Hund's rule is not obeyed. The orbitals in the $2p$ sublevel have the same energy. We must place electrons into individual $2p$ orbitals singly until an electron occupies all three orbitals before starting to pair electrons. Since there are four electrons in the $2p$ sublevel, two should be paired and two should be unpaired:

	$1s$	$2s$	$2p$		
O	↑↓	↑↓	↑↓	↑	↑

(c) The orbital diagram for magnesium is not correct. The two electrons in the $3s$ orbital have the same spin, so the Pauli exclusion principle is not followed. The arrows in the $3s$ orbital should be pointing in opposite directions to represent electrons with opposite spin:

	$1s$	$2s$	$2p$			$3s$
Mg	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓

Solution:

(a) The atomic number of sodium is 11, so it has 11 protons and 11 electrons. We distribute electrons into the orbitals in Figure 7.18, starting with the lowest-energy orbital, 1*s*. By putting a maximum of two electrons in each orbital, we get the electron configuration

$$1s^2 2s^2 2p^6 3s^1$$

(b) The atomic number of argon is 18. It has 18 electrons, which is enough to fill sublevels through 3*p*. The electron configuration is

$$1s^2 2s^2 2p^6 3s^2 3p^6$$

(c) From the periodic table we see that bromine has 35 electrons. Filling these electrons into the orbital diagram from lowest- to highest-energy sublevels, we get the following electron configuration:

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

Electrons fill the 4*s* orbital before filling the 3*d* orbitals because the 4*s* sublevel is lower in energy than the 3*d*.

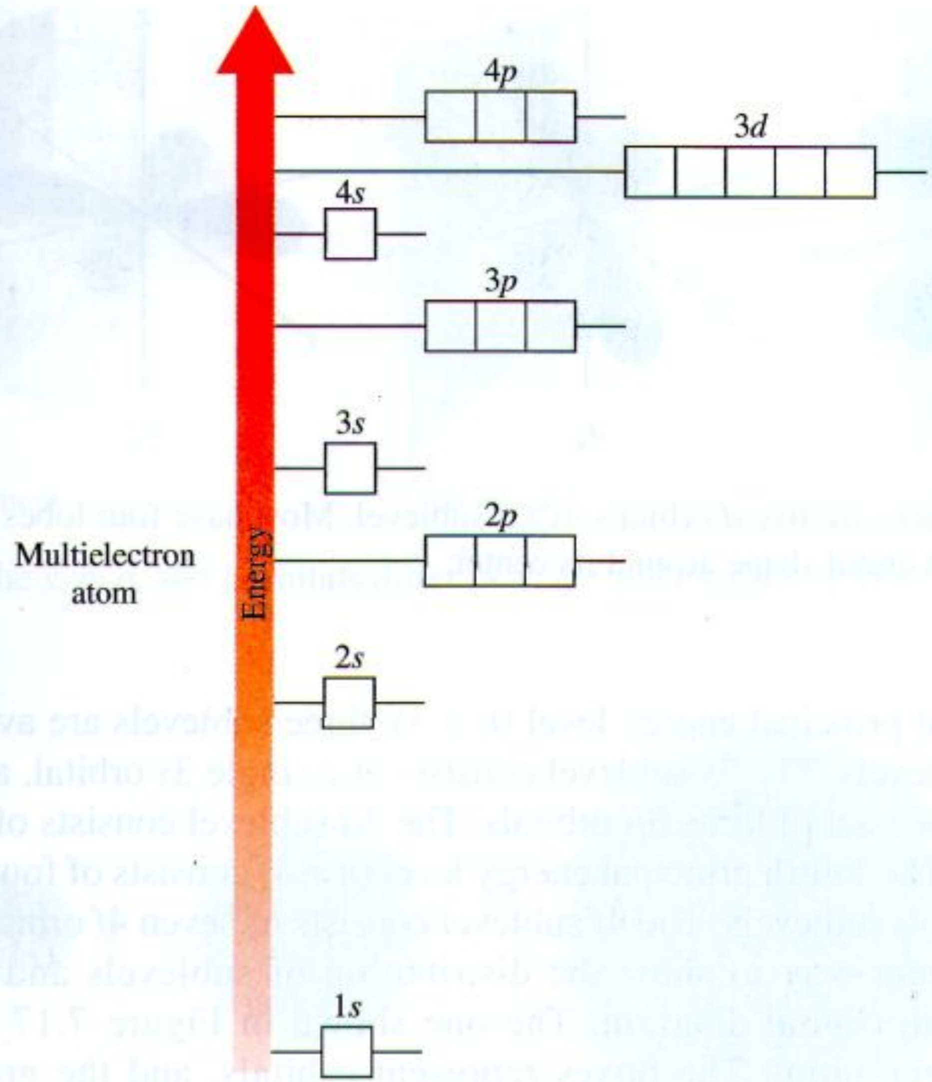


FIGURE 7.18 The orbital diagram for a multielectron atom shows the same sublevels and orbitals as for the hydrogen atom. However, in atoms that contain more than one electron, the sublevels within a principal energy level have different energies. Note that the energy of the 3*d* sublevel is higher than the energy of the 4*s* sublevel.

9.1 $\cdot\mathrm{Ba}\cdot + 2\cdot\mathrm{H} \longrightarrow \mathrm{Ba}^{2+}\ 2\mathrm{H}:^{-}$ (or BaH_2)

$[\mathrm{Xe}]6s^2 \quad 1s^1 \qquad [\mathrm{Xe}] \quad [\mathrm{He}]$

Example 9.2

DCM, Worked Examples

Classify the following bonds as ionic, polar covalent, or covalent: (a) the bond in HCl, (b) the bond in KF, and (c) the CC bond in H_3CCH_3.

Strategy We follow the 2.0 rule of electronegativity difference and look up the values in Figure 9.5.

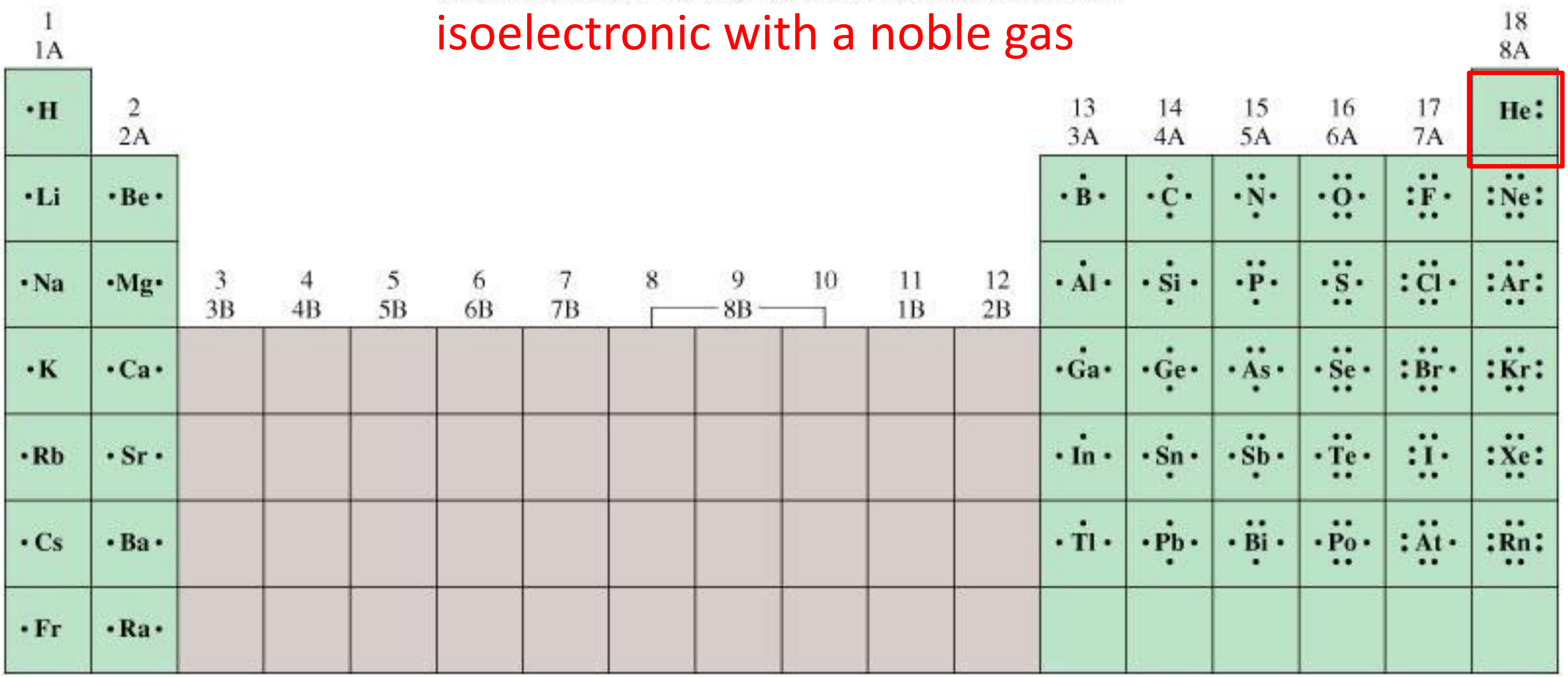
Solution (a) The electronegativity difference between H and Cl is 0.9, which is appreciable but not large enough (by the 2.0 rule) to qualify HCl as an ionic compound. Therefore, the bond between H and Cl is polar covalent.

(b) The electronegativity difference between K and F is 3.2, which is well above the 2.0 mark; therefore, the bond between K and F is ionic.

(c) The two C atoms are identical in every respect—they are bonded to each other and each is bonded to three other H atoms. Therefore, the bond between them is purely covalent.

Lewis Dot Symbols for the Representative Elements & Noble Gases

- maximum stability results when an atom is isoelectronic with a noble gas



- consists of the symbol of an element and one dot for each valence electron in an atom of the element