

# BIG IDEA 1: ATOMIC STRUCTURE AND PERIODICITY

## Big Idea 1

The chemical elements are fundamental building materials of matter, and all matter can be understood in terms of arrangements of atoms. These atoms retain their identity in chemical reactions.

Atoms are the fundamental building blocks of the elements which comprise all matter. Big Idea 1 relates to 3 main topics in this chapter: atomic structure, periodicity, and macroscale matter. The structure of the atom is the basis for the organization of the elements on the periodic table. Physical and chemical changes that we see on the macroscopic scale are the result of how atoms interact with other atoms and with electromagnetic radiation. Because atoms are so incredibly small we use moles as a unit of measurement when describing quantities of atoms.

You should be able to

- Use evidence to explain why atomic models have changed throughout history.
- Calculate the average atomic mass of an atom from mass spectrometry and isotopic data.
- Use photoelectron spectroscopy as evidence to show that electrons exist in quantized shells.
- Justify the arrangement of the periodic table based on the quantum mechanical model of the atom.

- Convert between mass, moles, and numbers of representative particles (atoms, molecules, formula units) of a substance.
- Calculate the molar mass of a compound and the percent composition of an element in a compound.
- Express the law of conservation of mass using symbolic representations and particulate models.

### AP Tip

Writing coherent answers to the free-response questions is essential. You need to differentiate between restating facts and explaining the concepts and evidence that supports the facts. Questions that involve mathematical problem solving should clearly show your work and answer. Partial credit may be awarded for some questions.

## PART 1: ATOMIC STRUCTURE

### EVIDENCE FOR ATOMS

#### MODELS CHANGE OVER TIME

Atomic models have been around since Democritus first thought about dividing matter by continually cutting it into smaller and smaller pieces until it could no longer be cut: *atomos*. The idea of atoms was opposed by Aristotle and it took about 2500 years until scientific evidence was used to revive the idea of atoms.

### HISTORY OF THE ATOM

(Chemistry 8th ed. pages 42–50/9th ed. pages 44–54)

#### JOHN DALTON

Dalton's model of the atom was based on the work of many scientists. He used the law of definite proportion and the law of multiple proportions to provide evidence for his model.

Dalton's model has 4 parts:

1. Each element is made of atoms.
2. Atoms of a given element are identical; atoms of different atoms are different in some fundamental ways.
3. Atoms can combine to form compounds and a given compound always has the same relative number and types of atoms.
4. Chemical reactions involve rearrangement of the ways the atoms are bonded together. The atoms themselves are not changed.

(Chemistry 8th ed. pages 41–44/9th ed. pages 44–47)

## THE LAW OF DEFINITE PROPORTION

The law of definite proportion states that a given compound always contains exactly the same proportions of elements by mass.

**EXAMPLE:** A sample of  $\text{H}_2\text{SO}_4$  contains 2.02 g hydrogen, 32.07 g sulfur, and 64.00 g oxygen. How many grams of sulfur and grams of oxygen are present in a second sample of  $\text{H}_2\text{SO}_4$  containing 7.27 g of hydrogen?

**SOLUTION:** Hydrogen is increased by a factor of 3.60 ( $7.27/2.02$ ). Therefore

$$\text{g sulfur} = 32.07 \times 3.60 = 115.45 \text{ g sulfur} = 115 \text{ g sulfur}$$

$$\text{g oxygen} = 64.00 \text{ g} \times 3.60 = 230.4 \text{ g oxygen} = 230 \text{ g oxygen}$$

## THE LAW OF MULTIPLE PROPORTIONS

The law of multiple proportions: When two elements form a series of compounds, the ratio of masses of the second element that combine with 1 gram of the first element can always be reduced to the smallest whole numbers.

**EXAMPLE:** Sulfur and oxygen can react to form both sulfur dioxide and sulfur trioxide. In sulfur dioxide there are 32.06 g sulfur and 32.00 g oxygen. In sulfur trioxide there are 32.06 g sulfur combined with 48.00 g oxygen. What is the ratio of the weights of oxygen that combine with 32.06 g sulfur?

**SOLUTION:**  $48.00/32.00 = 1.5$  or 3:2, demonstrating the law of multiple proportions.

## DISCOVERY OF THE NUCLEUS

Dalton's model of hard spheres that were identical for each element was quickly challenged because the model could not account for the way electricity interacted with matter and also did not account for the fact that most elements were composed of isotopes.

## ISOTOPES

(Chemistry 8th ed. pages 78–81/9th ed. pages 83–85)

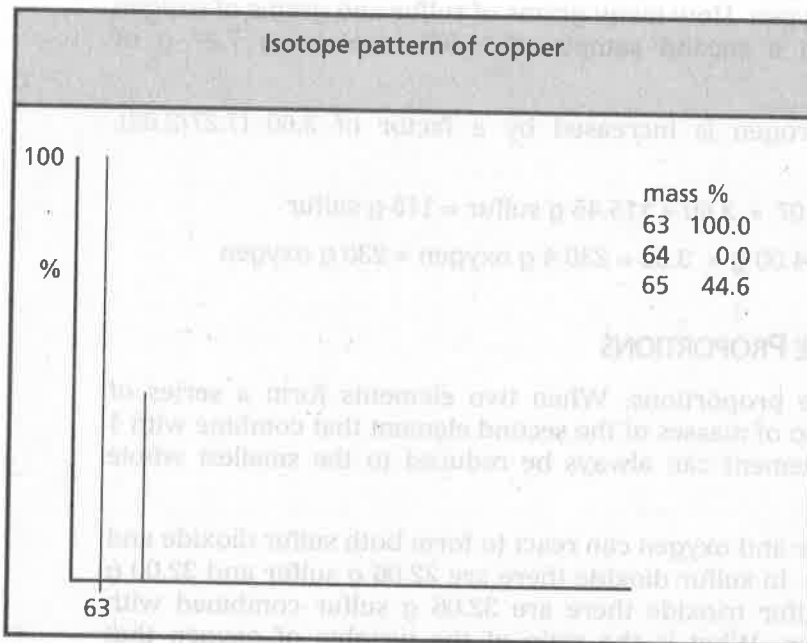
Isotopes are atoms of the same element with different numbers of neutrons and therefore different atomic masses.

## MASS SPECTROMETRY

(Chemistry 8th ed. pages 78–80/9th ed. pages 83–85)

Mass spectrometry is an analytical technique that is used to determine the mass spectrum of an element. Most elements exist as isotopes of different mass, and the mass spectrometer can determine the ratio of the masses.

An element sample is ionized and the resulting fragments are passed through a magnetic field that deflects the ions based on mass and charge. The deflection separates the ions and a unique spectrum is created. The relative heights of peaks on the spectrum provide evidence for the isotopic ratio of the element.



Isotope	Atomic Mass (amu)
$^{63}\text{Cu}$	62.9
$^{65}\text{Cu}$	64.9

The mass spectrum of copper is shown above. The highest peak is often set to 100%, but that is just an arbitrary scale. The relationship between the peaks is what is important. If you had 1446 atoms of copper, the data show that 1000 of those atoms would be copper-63 and 446 atoms would be copper-65. An average atomic mass could be calculated from these data.

$$\frac{[(1000 \text{ atoms})(62.9 \frac{\text{amu}}{\text{atom}}) + (446 \text{ atoms})(64.9 \frac{\text{amu}}{\text{atom}})]}{1446 \text{ atoms}} = 63.5 \text{ amu}$$

You are probably more familiar with calculating average atomic mass from percent abundance data, but the percent composition data originally came from mass spectrometry experiments. In this case, 69.16% of the sample (1000/1446 atoms) was copper-63 and the remaining 30.84% of the sample (446/1446 atoms) was copper-65.

## CALCULATION OF AVERAGE ATOMIC MASS FROM ISOTOPIC DATA

(Chemistry 8th ed. pages 78–81/9th ed. pages 83–85)

The average atomic mass of an element can be calculated from the percent abundance and mass of each isotope for that element.

Avg Atomic Mass = (percent abundance of isotope × mass number)

**EXAMPLE:** Element “E” is present with the following mass values and natural abundances.

Isotope	Mass Number (amu)	Percent Abundance (%)
$^{10}\text{E}$	10.01	19.78
$^{11}\text{E}$	11.01	80.22

What is the average atomic mass of the element, E? What is the element?

**SOLUTION:**  $(0.1978)(10.01) + (0.8022)(11.01) = 10.812$  amu.

The element is boron.

## NIELS BOHR

(Chemistry 8th ed. page 295–300/9th ed. pages 306–310)

Niels Bohr used the line spectrum of hydrogen to develop his atomic model. Bohr proposed that the electron in the hydrogen atom moves around the nucleus only in certain allowed circular orbits and the line spectra of elements are the result of electrons moving between these allowed orbits. Bohr’s model has often been described as a solar system model. Since a moving electron radiates energy, this model was unstable using classical mechanics and the model was shown to be incorrect. Furthermore, the model only worked for hydrogen and could not predict the spectra of other elements. An understanding of the electromagnetic spectrum is necessary to understand Bohr’s model and subsequent atomic models.

## ELECTROMAGNETIC RADIATION

(Chemistry 8th ed. pages 285–286/9th ed. pages 296–298)

Electromagnetic radiation provides an important means of energy transfer. Light from the sun reaches the earth in the form of radiation that includes ultraviolet and visible radiation. Electromagnetic radiation, from the shortest to the longest wavelength, includes gamma rays, x-rays, ultraviolet, visible, infrared, microwaves, and radio waves. All types of electromagnetic radiation exhibit both wave-like and particle-like behavior and travel at the speed of light in a vacuum.

Chemists have used the interaction of light with matter to study the structure of the atom.

## CHARACTERISTICS OF WAVES

(Chemistry 8th ed. pages 285–286/9th ed. pages 296–298)

Wavelength,  $\lambda$ , is the distance between two consecutive peaks or troughs in a wave.

Frequency,  $\nu$ , is the number of waves (cycles) per second that passes a given point in space.

Wavelength and frequency are related by the equation:

$$\nu = c / \lambda$$

in which  $c$  is the speed of light in a vacuum,  $2.9979 \times 10^8$  m/s. Wavelength is measured in meters, and frequency is measured in 1/s, or hertz (Hz).

## ENERGY OF ELECTROMAGNETIC RADIATION

(Chemistry 8th ed. pages 285–287/9th ed. pages 296–298)

Energy is quantized and can occur only in whole number multiples of  $h\nu$ . Each small packet of energy is called a quantum. A system can transfer energy only in discrete quanta. The change in energy for a system can be represented by

$$\Delta E = nh\nu$$

where  $n$  is an integer (1,2,3,...),  $h$  is Planck's constant which equals  $6.626 \times 10^{-34}$  Js, and  $\nu$  is the frequency.

Einstein viewed electromagnetic energy as a stream of particles called photons. And the energy of each photon can be given by

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

**EXAMPLE:** The laser in an audio compact disc player uses light with a wavelength of  $7.80 \times 10^2$  nm. Calculate the frequency of this radiation. What is the energy of this radiation per photon?

**SOLUTION:**  $1 \text{ nm} = 10^{-9} \text{ m}$

$$(7.80 \times 10^2 \text{ nm})(1 \times 10^{-9} \text{ m}/1 \text{ nm}) = 7.80 \times 10^{-7} \text{ m}$$

$$\nu = \frac{c}{\lambda} = \frac{2.998 \times 10^8 \text{ m/s}}{7.80 \times 10^{-7} \text{ m}} = 3.84 \times 10^{14} \text{ s}^{-1}$$

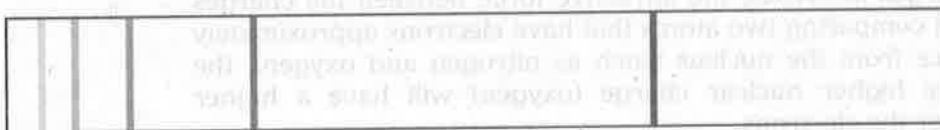
$$\begin{aligned} E_{\text{photon}} &= hc/\lambda \\ &= (6.626 \times 10^{-34} \text{ Js})(2.998 \times 10^8 \text{ m/s}) / 7.80 \times 10^{-7} \text{ m} \\ &= 2.55 \times 10^{-19} \text{ J} \end{aligned}$$

## ATOMIC SPECTRUM

(Chemistry 8th ed. pages 294–295/9th ed. pages 305–306)

A continuous spectrum contains all wavelengths of visible light. A line spectrum contains only a few lines, each corresponding to a discrete wavelength. When atoms are excited by adding energy, electrons move to higher energy levels as they absorb the energy. When the electrons return to their ground-state energies, the excess energy is given off by emitting light of various wavelengths to produce an emission spectrum or line spectrum.

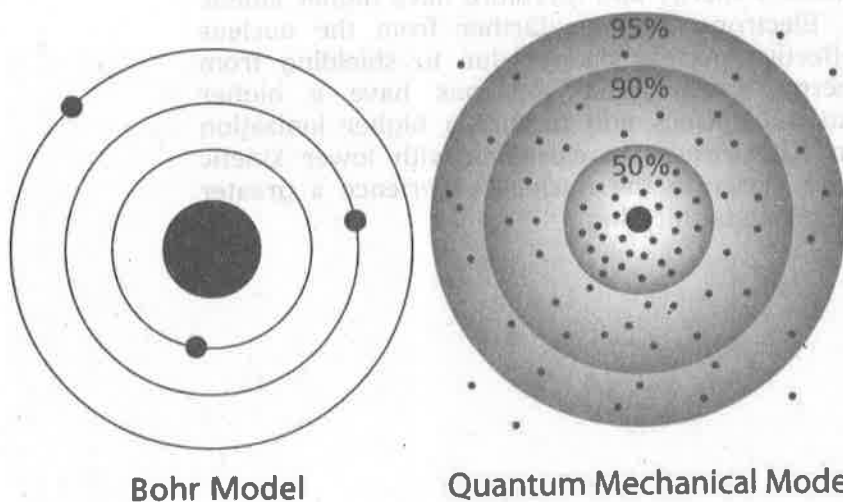
**The Emission lines for the element Hydrogen which correspond to the absorption lines for the same element**



## QUANTUM MECHANICAL MODEL

(Chemistry 8th ed. pages 300–303/9th ed. pages 310–313)

Our current model of the atom is the quantum mechanical model. In this model, we treat the electrons as a wave and do not believe it is possible to describe the path of an electron. An orbital is a three-dimensional electron density map in which there is a very high probability of finding the electron. The Heisenberg uncertainty principle is an important part of the model. It states that we cannot know both the momentum and location of an electron. The quantum mechanical (QM) model is an improvement over the Bohr model because the QM model supports evidence of energy levels and sublevels within atoms. The QM model can be approximated using computers and is the basis for software that predicts molecular shape and reactivity.



## ELECTROSTATIC FORCES HOLD THE ATOM TOGETHER

A neutral atom is composed of protons and neutrons in the nucleus and electrons in energy levels or “shells” outside of the nucleus. Coulomb’s law states that the force between the two particles is proportional to the magnitude and inversely proportional to the square of the distance between them.

$$|F| = k_e \frac{|q_1 q_2|}{r^2}$$

where  $F$  = force,  $k_e$  = proportionality constant,  $q$  = charges on each particle and  $r$  is the distance between the particles.

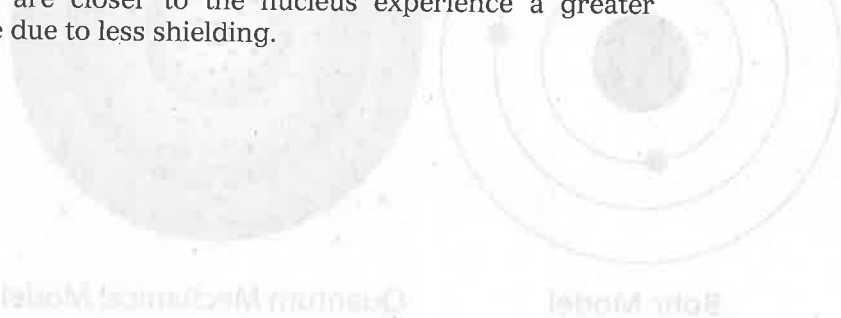
If the charges are of opposite signs, the force is attractive and if the charges are of the same sign, the force is repulsive. As the distance between the charges increases, the attractive force between the charges decreases. When comparing two atoms that have electrons approximately the same distance from the nucleus (such as nitrogen and oxygen), the nucleus with the higher nuclear charge (oxygen) will have a higher attractive force for the electrons.

## PHOTOELECTRON SPECTROSCOPY

Photoelectron spectroscopy (PES) is a technique that can be used to provide evidence for electron shells in an atom. Light is shined on an atom and the minimum frequency needed to remove an electron is determined. If the frequency of light is greater than the threshold needed to remove an electron, the extra energy is retained by the ionized atom as kinetic energy. Because energy is conserved, the energy of light that hits the atom ( $h$ ) must equal the energy required to remove the electron (IE) and the kinetic energy (KE) of the removed electron.

$$h = KE + IE$$

Since the frequency of the light hitting the atom is known and the kinetic energy of the removed electron can be measured, the ionization energy of the electron can be calculated. Electrons that are farther away from the nucleus require less ionization energy and therefore have higher kinetic energy when removed. Electrons that are farther from the nucleus experience a smaller effective nuclear charge due to shielding from electrons in lower energy levels. Inner electrons have a higher electrostatic attraction to the nucleus and require a higher ionization energy to remove them. This results in electrons with lower kinetic energy. Electrons that are closer to the nucleus experience a greater effective nuclear charge due to less shielding.

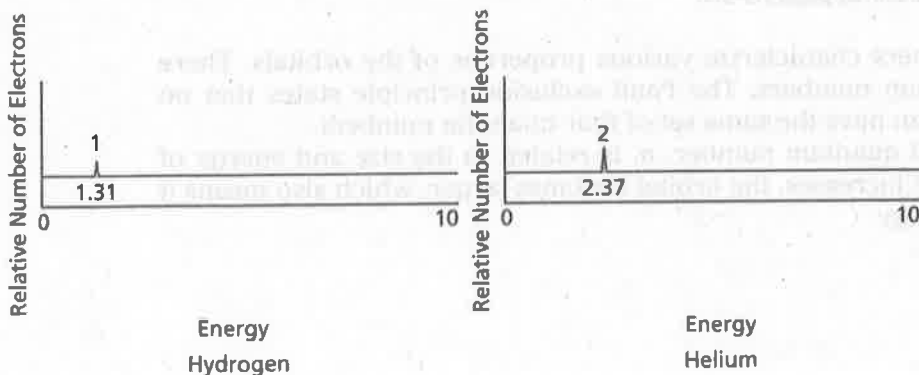




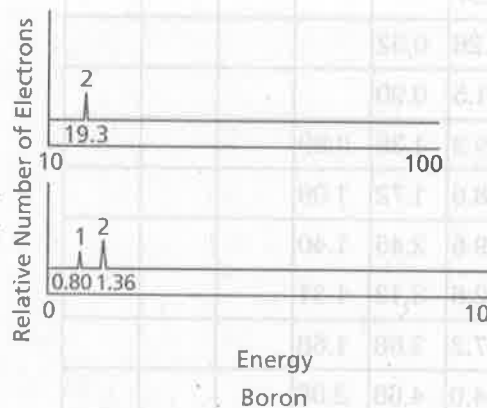
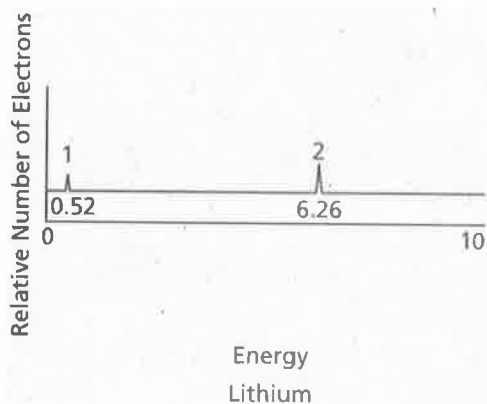
The ionization energies of the first twenty elements are listed in the table below.

Element	Ionization Energy (MJ/mol)					
H	1.31					
He	2.37					
Li	6.26	0.52				
Be	11.5	0.90				
B	19.3	1.36	0.80			
C	28.6	1.72	1.09			
N	39.6	2.45	1.40			
O	52.6	3.12	1.31			
F	67.2	3.88	1.68			
Ne	84.0	4.68	2.08			
Na	104	6.84	3.67	0.50		
Mg	126	9.07	5.31	0.74		
Al	151	12.1	7.79	1.09	0.58	
Si	178	15.1	10.3	1.46	0.79	
P	208	18.7	13.5	1.95	1.01	
S	239	22.7	16.5	2.05	1.00	
Cl	273	26.8	20.2	2.44	1.25	
Ar	309	31.5	24.1	2.82	1.52	
K	347	37.1	29.1	3.93	2.38	0.42
Ca	390	42.7	34.0	4.65	2.9	0.59

Both hydrogen and helium have a single ionization energy that corresponds to the  $n = 1$  energy level electrons. Even though helium has 2 electrons, they are both approximately equally distant from the nucleus and therefore have the same ionization energy.



Lithium has 2 ionization energies that correspond to  $n = 1$  and  $n = 2$  energy levels. Boron's ionization energies, however, do not seem to follow the pattern at first. The IE for  $n = 1$  (19.3 MJ/mol) is about 14 times greater than  $n = 2$  (1.36 MJ/mol). There is less than 2 times difference between the two sets of outermost electrons (1.36 MJ/mol and 0.80 MJ/mol). This evidence supports the idea of 2 subshells in  $n = 2$ . The subshell closer to the nucleus is called 2s and the one farther away is called 2p.



**EXAMPLE:** How do the data in the table above support the claim that calcium's electrons are in 4 shells?

**SOLUTION:** The first ionization energy of calcium is about 9 times larger than the second indicating that electrons are in  $n = 1$ . The second and third ionization energies are about 1.25 times different indicating that those electrons are probably in the same shell ( $n = 2$ ). The 3rd and 4th energy level are approximately 7 times different indicating the existence of another shell ( $n = 3$ ).  $IE_4$  is about 1.6 times larger than  $IE_5$  indicating 2 subshells in  $n = 3$ . Finally,  $IE_5$  is about 5 times larger than  $IE_4$  indicating another energy level,  $n = 4$ .

The core electrons effectively shield the valence electrons from the positive charge of the nucleus due to the repulsion of like charges. Less energy is required to remove electrons and overcome the attractive forces of the nucleus. This phenomenon also explains the differences in ionization energy in subshells of the same energy level.

## QUANTUM NUMBERS

(Chemistry 8th ed. pages 303–304/9th ed. pages 313–314)

Quantum numbers characterize various properties of the orbitals. There are four quantum numbers. The Pauli exclusion principle states that no two electrons can have the same set of four quantum numbers.

The principal quantum number,  $n$ , is related to the size and energy of the orbital. As  $n$  increases, the orbital becomes larger, which also means it is higher in energy.

The angular momentum quantum number,  $\ell$ , is related to the shape of atomic orbitals.

The magnetic quantum number,  $m_\ell$  is related to the orientation of the orbital in space relative to the other orbitals in the atom.

The electron spin quantum number,  $m_s$ , has only two values:  $+1/2$  and  $-1/2$ . This quantum number is necessary to describe the fact that the electron has two possible orientations, or spins, when placed in an external magnetic field. Because electrons can only adopt two different spins, each orbital can hold a maximum of two electrons.

The four quantum numbers are summarized in the following table.

<b>Quantum Numbers</b>		
<b>Name</b>	<b>Designation</b>	<b>Property of the Orbital</b>
Principal quantum number	$n$	Related to size and energy of the orbital
Angular momentum quantum number	$\ell$	Related to the shape of the orbital
Magnetic quantum number	$m_\ell$	Related to the position of the orbital in space in relation to other orbitals
Electron spin quantum number	$m_s$	Related to the spin of the electron, which can be only one of two values

The principal quantum number,  $n$ , is the number which appears in front of the types of orbital  $s$ ,  $p$ ,  $d$ , or  $f$ . In general, for periods (rows) 1, 2, and 3, the principal quantum number,  $n$ , is the same as the row number. For periods 3, 4, 5, 6, and 7, the principal quantum number written before  $s$  and  $p$  orbitals is the same as the row number for the main group of elements, those in columns 1–2 and 13–18 (1A through 8A). For the transition elements, the value of  $n$  which is written in front of the  $d$  orbital is 1 number less than the row number. For example, scandium, in period 4, has a value of  $n$  equal to 3 for its  $d$  electron. In the sixth row of the periodic table, the  $f$ -transition elements occur after lanthanum. For these elements, the value of  $n$  which is placed before the  $f$  is 2 less than the row in which they occur.

## ORBITAL SHAPES AND ENERGIES

(Chemistry 8th ed. pages 305–308/9th ed. pages 314–318)

### s ORBITALS

The  $s$  orbitals have a spherical distribution of electron density. The  $s$  orbitals become larger as the value of  $n$  increases.

Nodes are areas in which there is zero probability of finding an electron. The  $2s$  has one node which separates areas of high probability, the  $3s$  has two nodes, and so forth.

### p ORBITALS

There are three  $p$  orbitals which first appear in  $n = 2$ . The  $p$  orbitals have two lobes which are separated by a node at the nucleus. The  $p$  orbitals are labeled  $n_{px}$ ,  $n_{py}$ , and  $n_{pz}$ , according to the axis along which the lobe lies.

Degenerate orbitals are orbitals with the same value of  $n$  which have the same energy. The three  $p$  orbitals in the same energy level are degenerate or equal in energy.

### d AND f ORBITALS

The  $d$  orbitals first occur when  $n = 3$  but in the fourth period of the periodic table. The  $f$  orbitals first occur when  $n = 4$  and in the sixth period.

**EXAMPLE:** Which of the following orbital designations is incorrect:  $1s$ ,  $1p$ ,  $2d$ ,  $4f$ ?

**SOLUTION:**  $1p$  is incorrect:  $p$  orbitals do not exist in energy level 1;  $n = 2$  is the first energy level that has  $p$  orbitals

$2d$  is incorrect:  $d$  orbitals do not exist in energy level 2; the first energy level with  $d$  orbitals is energy level 3

**EXAMPLE:** Give the maximum number of electrons in an atom that can have the quantum number  $n = 3$ .

**SOLUTION:** Energy level 3 is comprised of 3 types of orbitals:  $s$ ,  $p$ , and  $d$ . Two electrons can exist in an  $s$  orbital, 6 in a  $p$  orbital, and 10 in a  $d$  orbital. Therefore, 18 electrons can exist in level 3. For any energy level,  $2n^2 = \text{max number of electrons for any energy level}$ .

**EXAMPLE:** Give the maximum number of electrons in an atom that can exist in  $2p$  orbitals.

**SOLUTION:** Each  $2p$  orbital can hold 2 electrons with spins in opposite directions. Since three orientations of  $p$  orbitals exist, a total of 6 electrons can be in  $2p$ .

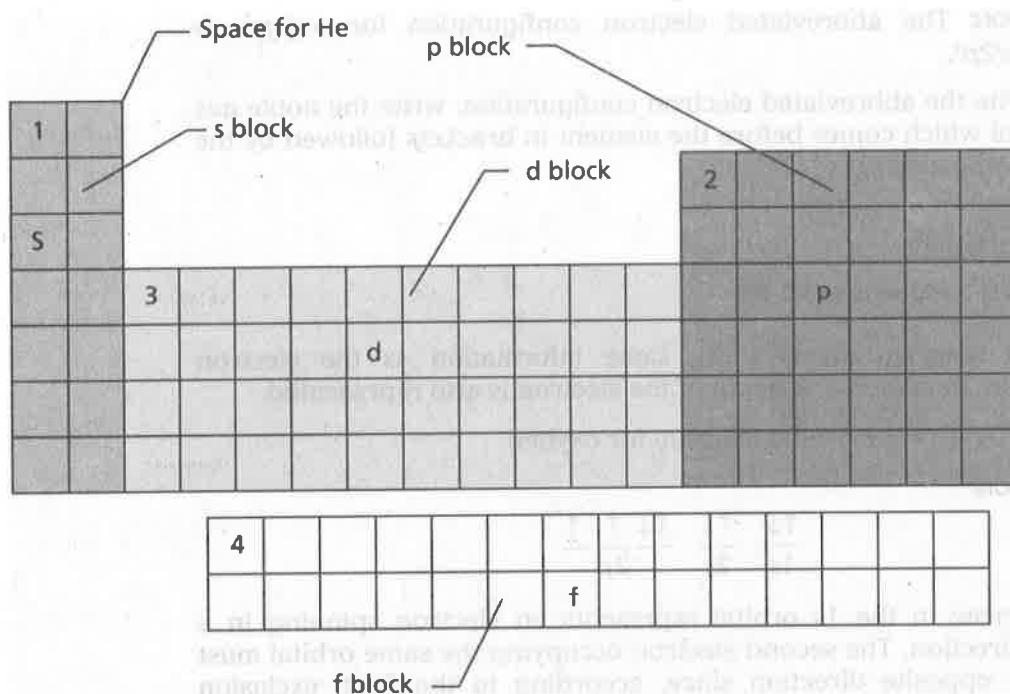
## PART 2: PERIODICITY

### ARRANGING THE ELEMENTS

#### ELECTRONIC STRUCTURE AND PERIODIC TABLE

(Chemistry 8th ed. pages 317–318/9th ed. pages 328–329)

The current arrangement of the periodic table is a result of the repeating pattern of electron arrangement as the atomic number of the elements increases.

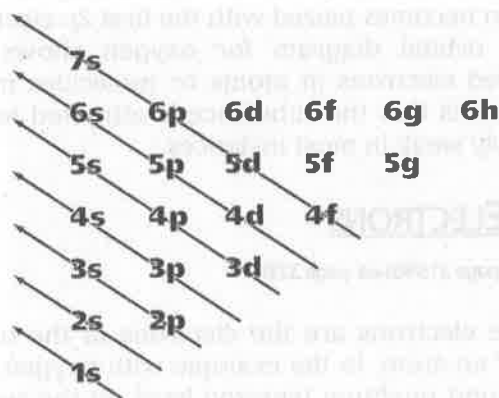


#### AUFBAU PRINCIPLE AND ELECTRON CONFIGURATIONS

(Chemistry 8th ed. pages 312–318/9th ed. pages 322–329)

The aufbau principle is a scheme used to reproduce the electron configuration of the ground states of atoms by successively filling sublevels with electrons in a specific order that is directly related to the amount of energy within each orbital.

The electron configuration for an atom of an element represents all of the electrons in the atom and it shows in which energy levels and orbitals the electrons reside.



**EXAMPLE:** Write an electron configuration for oxygen.

**SOLUTION:** The electron configuration for oxygen is  $1s^2 2s^2 2p^4$ .

To write an electron configuration for an atom of an element, follow the aufbau principle and add electrons to the lowest energy orbital, the  $1s$ . Using the Pauli exclusion principle, each orbital holds two electrons, we fill up the  $1s$  with two electrons of opposite spin, then the  $2s$ , to get  $1s^2 2s^2$ . We have used four electrons so far and oxygen has eight total electrons. The remaining four electrons go into the three  $2p$  orbitals. Hund's rule requires that we place one electron in each separate  $p$  orbital with parallel spin before pairing electrons. Write the electron configuration as  $1s^2 2s^2 2p^4$ .

**EXAMPLE:** Write the abbreviated electron configuration for oxygen.

**SOLUTION:** The abbreviated electron configuration for oxygen is  $[\text{He}] 2s^2 2p^4$ .

To write the abbreviated electron configuration, write the noble gas symbol which comes before the element in brackets followed by the remaining electrons.

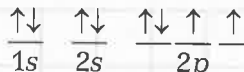
## ORBITAL DIAGRAMS

(Chemistry 8th ed. pages 312–313/9th ed. pages 322–323)

The orbital diagram displays the same information as the electron configuration. In addition, the spin of the electron is also represented.

**EXAMPLE:** Give the orbital diagram for oxygen.

**SOLUTION:**



The first arrow in the  $1s$  orbital represents an electron spinning in a particular direction. The second electron occupying the same orbital must spin in the opposite direction since, according to the Pauli exclusion principle, only two electrons with opposite spin can occupy the same orbital. When the  $1s$  orbital is filled, the  $2s$  orbital fills in the same manner as the  $1s$ . There are three degenerate  $2p$  orbitals, so one electron must go into each of the  $2p$  orbitals according to Hund's rule. The fourth  $2p$  electron becomes paired with the first  $2p$  electron.

The orbital diagram for oxygen shows unpaired electrons in  $2p$ . Unpaired electrons in atoms or molecules make particles paramagnetic. This means that the substance is attracted to a magnet. The attraction is relatively weak in most instances.

## VALENCE ELECTRONS

(Chemistry 8th ed. page 313/9th ed. page 323)

Valence electrons are the electrons in the outermost principal quantum level of an atom. In the example with oxygen, the valence electrons are in the second quantum (energy) level. In the second energy level, there are six valence electrons, two in the  $2s$  and four in the  $2p$ . Coincidentally, the valence electrons of a main group element are the same as the group

number in the old A-B group system, except for helium which has two valence electrons. All of the electrons that are not in the highest energy level are known as core electrons. The 1s electrons in oxygen are core electrons. Pages 315 and 316 of the 8th edition and pages 326 and 327 of the 9th edition of *Chemistry* display helpful diagrams which show how the orbitals are filled in various parts of the periodic table.

## PERIODIC TRENDS IN ATOMIC PROPERTIES

(*Chemistry* 8th ed. pages 318–323/9th ed. pages 329–334)

### ATOMIC RADII

(*Chemistry* 8th ed. pages 322–323/9th ed. pages 333–334)

Atomic radii decrease in going from left to right across a period from the alkali metals to the halogens. This can be explained by an increasing effective nuclear charge, which causes the electrons to be drawn closer to the nucleus.

Atomic radius increases down a group because of the increases in the orbital sizes in successive principal quantum levels. Also, there is an increase in the shielding of the nucleus by the core electrons, which decreases the effective nuclear charge felt by the valence electrons.

**EXAMPLE:** Arrange the following atoms in order of increasing atomic radius.

C, Al, F, Si

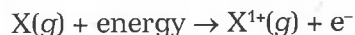
**SOLUTION:**  $F < C < Si < Al$ ; F and C are in the second period, which means that they have two shells or energy levels. Within a period, the radius decreases as the atomic number increases because the higher nuclear charge that results from added protons increases the attraction between the electrons and the nucleus. This means F will be smaller than C. Al and Si are in the third period and have one more energy level, so they will have a larger radius than atoms in period two. Since Si has more protons than Al, it has a smaller radius.

### IONIZATION ENERGY

(*Chemistry* 8th ed. pages 318–321/9th ed. pages 329–332)

Ionization energy is the energy required to remove an electron from a gaseous atom or ion.

The first ionization energy is the energy required to remove the highest energy electron (the one bound least tightly) which is one of the valence electrons.



For lithium, the first ionization energy is the energy required to remove an electron from the 2s orbital. The configuration for lithium is  $1s^22s^1$  and the valence electron is the 2s electron.

In general, the first ionization energy increases across a period from left to right. This trend can be related to the trend in atomic radii. The smaller the atom, the greater is its first ionization energy. In going across a period from left to right, atomic radii decrease as the charge on the nucleus increases. There are irregularities in this trend in ionization energy. The first ionization energy decreases from beryllium to boron because the electrons in the 2s orbital are much more effective at shielding the electrons in the 2p orbital than they are at shielding each other. In all atoms except for hydrogen, the np orbital is slightly higher in energy than the ns orbital. The higher energy 2p electron of boron is easier to remove than one of the 2s electrons of beryllium.

The first ionization energy decreases from nitrogen to oxygen due to repulsion of paired electrons in the  $p^4$  configuration of oxygen. Nitrogen has a valence electron configuration  $[\text{He}]2s^22p^3$  and oxygen has a valence electron configuration of  $[\text{He}]2s^22p^4$ .

The first ionization energy decreases down a group with increasing atomic number because the valence electrons become easier to remove as the atoms become larger going down a group. Shielding due to electron repulsion in core energy levels reduces the effective nuclear charge of the atoms. This lowers the attraction between the nucleus and the valence electrons which means less energy is needed to remove an electron.

The second ionization energy is the energy required to remove the second outermost electron. This value is larger than the first ionization energy. The first ionization energy removes an electron from a neutral atom. The second ionization energy removes an electron from a positive ion.



The increase in positive charge binds the electrons more tightly, and the ionization energy increases as each successive electron is removed. Also, a slightly larger increase in ionization energy occurs when removing an electron from a higher orbital, such as a 3p orbital vs. a 3s orbital. For example, the difference between the first and second ionization energies in magnesium is 709 kJ. Both of these electrons are 3s. The difference between the first and second ionization energy of aluminum, however, is 1238 kJ. The first electron removed is 3p which is effectively shielded by 3s. The second electron removed is 3s which has less shielding and therefore a stronger attraction to the nucleus. Note a large increase occurs in ionization energy when a core electron from a full energy level is removed.



## General increase

Element	Successive Ionization Energies for Period 3 Elements						
	IE <sub>1</sub>	IE <sub>2</sub>	IE <sub>3</sub>	IE <sub>4</sub>	IE <sub>5</sub>	IE <sub>6</sub>	IE <sub>7</sub>
Na	498	4560	6910	9540	13 400	16 600	20 100
Mg	736	1445	7730	10 600	13 600	18 000	21 700
Al	577	1815	2740	11 600	15 000	18 310	23 290
Si	787	1575	3220	4350	16 100	19 800	23 800
P	1063	1890	2905	4950	6270	21 200	25 400
S	1000	2260	3375	4565	6950	8490	27 000
Cl	1255	2295	3850	5160	6560	9360	11 000
Ar	1519	2665	3945	5770	7230	8780	12 000

General decrease

**EXAMPLE:** Explain why the difference between the third ionization energy and fourth ionization energy for aluminum is so large in comparison with the first and second and second and third ionization energies.

**SOLUTION:** Aluminum has the valence electron configuration [Ne]3s<sup>2</sup>3p<sup>1</sup>. The first electron removed is the 3p, the second electron removed is a 3s, and the third electron removed is the remaining 3s electron. The fourth electron, a core electron (a 2p), requires much more energy for removal than the valence electrons, the 3s and the 3p electrons, because it is more tightly bound due to a higher effective nuclear charge.

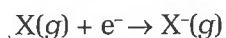
**EXAMPLE:** Arrange the following atoms in order of increasing first ionization energy. I, Rb, Na

**SOLUTION:** Na will have the smaller radius. Rb < I < Na. Na has the largest IE<sub>1</sub> because the electron is removed from the 2s sublevel. Both Rb and I have valence electrons in the 5th energy level, but I has 53 protons and a smaller radius than Rb with only 37 protons which makes I smaller than Rb. The smaller the atom, the more energy required to remove the valence electron.

## ELECTRON AFFINITY

(Chemistry 8th ed. pages 321–322/9th ed. pages 332–333)

Electron affinity is the energy change associated with the addition of an electron to a gaseous atom.



In general, electron affinity becomes increasingly negative or exothermic as we proceed across a period with increasing atomic number. The greater the charge on the nucleus of the atom, the greater is the affinity for electrons. The addition to a noble gas, however, would require

the added electron to reside in a new, higher energy level which is energetically unfavorable, so that stable, isolated  $X^-$  ions do not exist. Electron affinities are not measured for noble gases.

Electron affinities become slightly more positive, less exothermic, going down a group. The electron–nucleus attraction decreases down a group due to increasing atomic size and shielding. But the added electron experiences less electron–electron repulsion, since the orbital to which the electron is being added, is increasingly spread out. One exception to this trend is fluorine. The value of electron affinity for F is more positive than that of Cl due to the small size of the F atom and the greater electron–electron repulsion in the smaller  $2p$  sublevel.

**EXAMPLE:** Arrange the following elements from the least to the most exothermic electron affinity. Cl, Al, P and justify your answer.

**SOLUTION:** From the least to the most exothermic is Al, P, Cl. In general, the greater the effective nuclear charge on the atom, the more exothermic the electron affinity. Higher effective nuclear charges will have greater attraction between the positive nucleus and the negative electrons. Since all attractions release energy, a greater attraction will release more energy.

## IONIC RADII

(Chemistry 8th ed. pages 352–353/9th ed. pages 363–365)

Cations are smaller than their parent atoms and anions are larger than their parent atoms. The size of an ion depends on its nuclear charge, the number of electrons it contains, and the outer energy level of the atom. Positive ions are always smaller than neutral parent atoms because the nuclear charge is the same but the number of electrons has decreased, increasing the effective nuclear charge felt by each remaining electron. When all the valence electrons are lost, the change in volume is greater because as  $n$  decreases, the radial probability is smaller so the radius is smaller.

Anions are larger than their parent atoms because the electrons gained cause the proton to electron ratio to decrease, and the effective nuclear charge felt by each electron decreases. Electron–electron repulsions also increase. The electrons spread out more to accommodate the additional electrons.

In an isoelectronic series, ions contain the same number of electrons.

**EXAMPLE:** Arrange the following ions in order of increasing size:  $Ca^{2+}$ ,  $S^{2-}$ ,  $K^+$ ,  $Cl^-$ .

**SOLUTION:** All of the ions mentioned above are isoelectronic with argon, containing 18 electrons each. In order of increasing size  $Ca^{2+} < K^+ < Cl^- < S^{2-}$ .  $Ca^{2+}$  is the smallest because it contains the greatest number of protons. The greater the positive charge on the nucleus, the stronger the attraction for electrons.

## PART 3: MATTER ON THE MACRO SCALE

### THE MOLE

(Chemistry 8th ed. pages 81–84/9th ed. pages 85–90)

The mole is defined as the number of atoms in exactly 12 g of carbon-12. Calculations involving moles utilize the conversions in the table below.

1 mol of ...	is ...
monoatomic element	$6.022 \times 10^{23}$ atoms
molecular compound or diatomic element	$6.022 \times 10^{23}$ molecules
ionic compound	$6.022 \times 10^{23}$ formula units

### MOLAR MASS

(Chemistry 8th ed. pages 84–87/9th ed. pages 90–92)

Molar mass is the mass of 1 mole of an element or compound.

**EXAMPLE:** The principal ore in the production of aluminum cans has a molecular formula of  $\text{Al}_2\text{O}_3 \cdot 2\text{H}_2\text{O}$ . What is the mass in grams of  $2.10 \times 10^{24}$  formula units (f.u.) of  $\text{Al}_2\text{O}_3 \cdot 2\text{H}_2\text{O}$ ?

**SOLUTION:** The molar mass of  $\text{Al}_2\text{O}_3 \cdot 2\text{H}_2\text{O}$  is approximately 138 g/mol. Since this is an ionic compound, 1 mol contains  $6.022 \times 10^{23}$  formula units.

$$2.10 \times 10^{24} \text{ formula units} \times \frac{1 \text{ mol } \text{Al}_2\text{O}_3 \cdot 2\text{H}_2\text{O}}{6.022 \times 10^{23} \text{ f.u.}} \times \frac{138 \text{ g}}{1 \text{ mol } \text{Al}_2\text{O}_3 \cdot 2\text{H}_2\text{O}} = 481 \text{ g}$$

### PERCENT COMPOSITION

(Chemistry 8th ed. pages 88–90/9th ed. pages 94–96)

The mass percentages of elements in a compound can be obtained by comparing the mass of each element present in 1 mole of the compound to the total mass of 1 mole of compound.

$$\frac{\text{Mass of element in 1 mol of compound}}{\text{Mass of 1 mol of compound}} \times 100\%$$

**EXAMPLE:** Calculate the percent of oxygen in  $\text{Mg}(\text{NO}_3)_2$ .

**SOLUTION:** The molar mass of  $\text{Mg}(\text{NO}_3)_2$  is approximately 164 g/mol. There are 6 moles of oxygen in each mole of  $\text{Mg}(\text{NO}_3)_2$  so the molar mass of oxygen must be multiplied by 6.

$$\frac{\left(\frac{16 \text{ g O}}{\text{mol O}}\right) \left(\frac{6 \text{ mol O}}{\text{mol Mg}(\text{NO}_3)_2}\right)}{148 \text{ g mol}^{-1} \text{Mg}(\text{NO}_3)_2} \times 100 = 65\%$$

## CHANGES IN MATTER

Matter exists as pure substances or mixtures and can undergo physical or chemical changes.

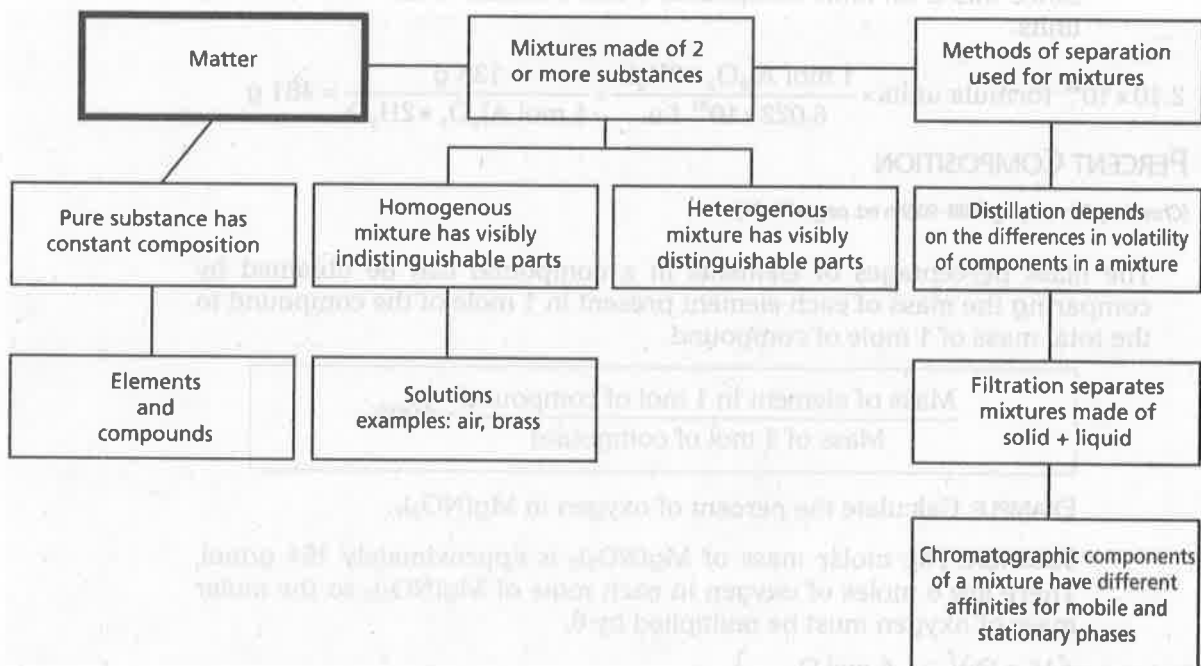
Physical changes in matter do not change the original composition of the substance. Changes in state such as boiling or melting are physical changes. Changes involving an alteration in the form of the substance such as grinding or tearing are physical. Physical properties are properties of a substance that can be observed without changing the composition of the substance. During a physical change, bonds are not broken and no reaction between atoms occurs. For example, density, color, and boiling point are physical properties.

Chemical changes in matter change the composition of the original substance by breaking and making bonds between atoms. A new substance is produced when a chemical change occurs. Evidence that a chemical change has occurred includes change in color or odor or the production of a gas or a solid (precipitate). Some examples of chemical properties include flammability and reactivity to air.

## MIXTURES AND PURE SUBSTANCES

(Chemistry 8th ed. pages 28–29/9th ed. pages 28–31)

Pure substances, such as elements or compounds, make up mixtures. The following chart summarizes mixtures and the methods for separating them.

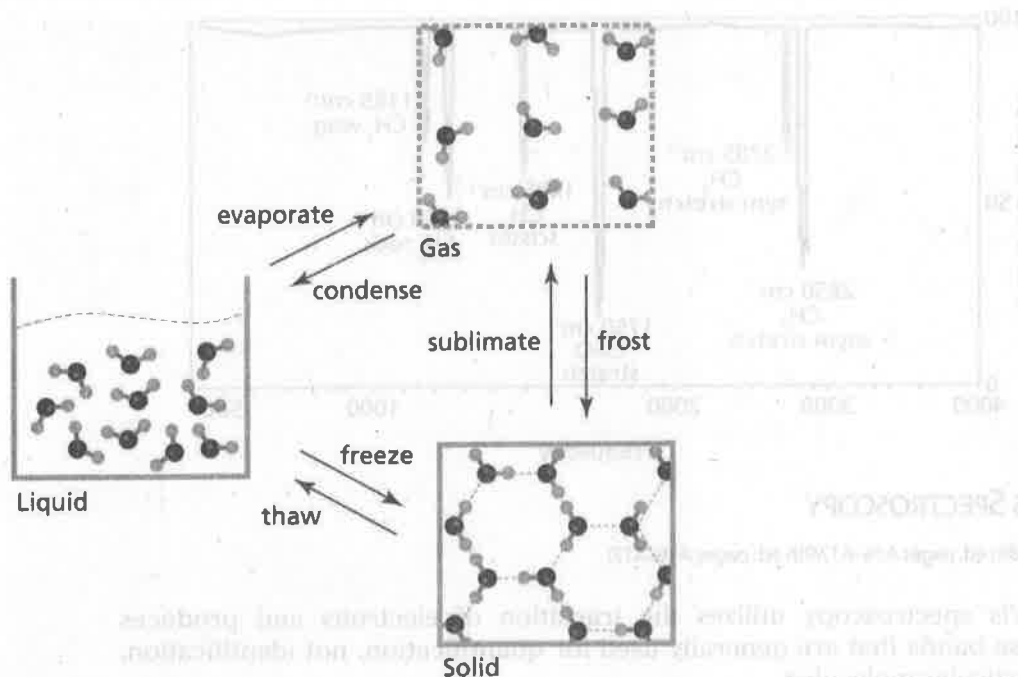


## CONSERVATION OF ATOMS

Matter cannot be created nor destroyed in any physical or chemical change. All changes of matter involve the same basic steps: energy is added to overcome attractions between the particles, the particles rearrange, and energy is released as new attractions are formed. Conservation of atoms is represented in particulate models and in symbolic representations that may include mathematical and/or chemical equations. Key examples of the recycling of matter include biogeochemical cycles such as the nitrogen cycle, carbon cycle, and water cycle.

The physical changes that occur as water goes through its phase changes are shown below.

### PARTICULATE MODEL



### SYMBOLIC REPRESENTATION



## CONSERVATION OF MASS

(Chemistry 8th ed. page 41/9th ed. page 44)

The number of particles (atoms, formula units, molecules) in a given mass of a substance can be calculated by knowing the chemical formula of the substance and applying the mole concept. The coefficients in a balanced chemical equation indicate the relative number of particles involved in the reaction. Since the number of atoms is constant, the total mass of reactants and products must also be constant.

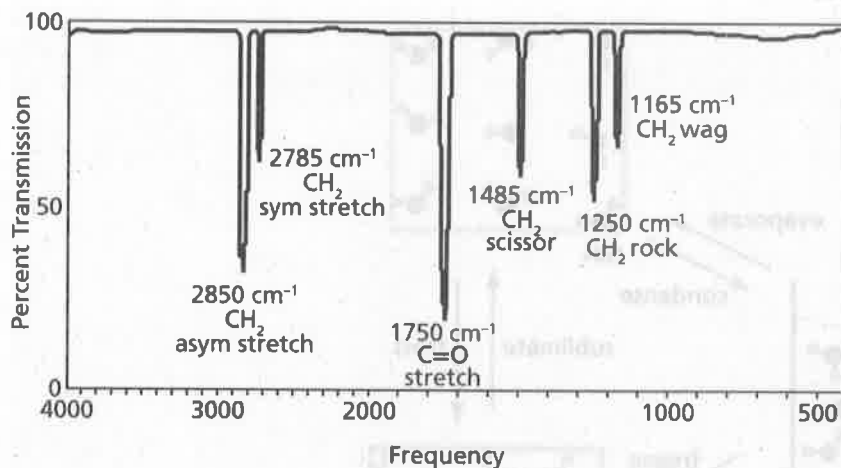
## IDENTIFYING AND MEASURING MATTER

(Chemistry 8th ed. pages A16–A19/9th ed. pages A16–A19)

Atoms and molecules interact with electromagnetic radiation in a variety of ways. The most commonly used types of spectral analysis include infrared spectroscopy (IR) and ultraviolet/visible spectroscopy (UV-Vis).

### INFRARED SPECTROSCOPY

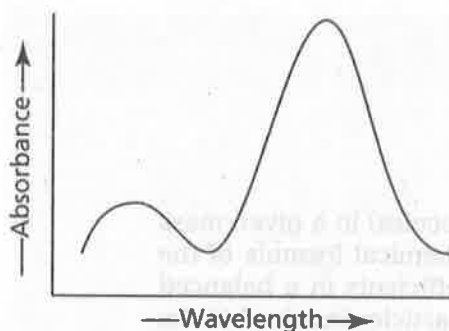
IR spectroscopy involves measuring the radiation absorbed by molecules as they stretch, bend, and twist. Certain functional groups such as  $\text{-OH}$  and  $\text{-C=O}$  have distinctive absorbance bands on IR spectra. Each molecule produces a unique spectrum that can be used for identification purposes.



### UV-VIS SPECTROSCOPY

(Chemistry 8th ed. pages A16–A17/9th ed. pages A16–A17)

UV-Vis spectroscopy utilizes the transition of electrons and produces diffuse bands that are generally used for quantification, not identification, of particular molecules.



**BEER'S LAW***(Chemistry 8th ed. pages A17–A19/9th ed. pages A17–A19)*

The amount of chemical can be quantified using Beer's law (Beer-Lambert law). The absorbance of an analyte at a given wavelength is directly proportional to the analyte concentration.

$$A = \epsilon \ell c$$

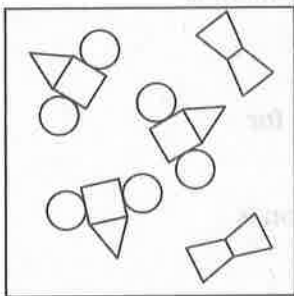
where  $A$  is the absorbance at the given wavelength,  $\epsilon$  is the molar absorptivity,  $\ell$  is the pathlength of the cell containing the chemical, and  $c$  is the concentration. Since  $\epsilon$  and  $\ell$  are constant, if the absorbance of a known concentration is measured, the concentration of an unknown substance can be determined from measuring the absorbance of the unknown.

**GRAVIMETRIC ANALYSIS AND TITRATION***(Chemistry 8th ed. pages 151–152, 157/9th ed. pages 160–161, 166)*

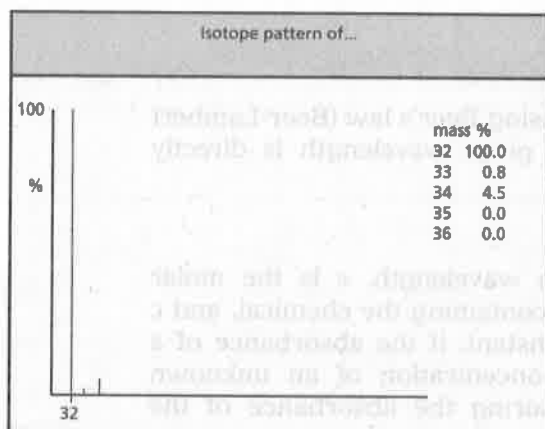
Analytical techniques such as gravimetric analysis and titration use conservation of mass to determine the quantity of an analyte in a solution. In gravimetric analysis the substance of interest is selectively precipitated from a solution and the mass of the precipitate is used to determine the mass of the analyte. A titration involves the use of an indicator to determine the stoichiometric equivalence point of a reaction. A common indicator is a color change. The point at which the change occurs is called the end point.

**MULTIPLE-CHOICE QUESTIONS**

No calculators are to be used in this section.

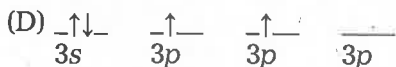
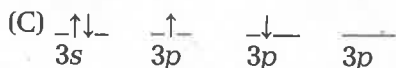
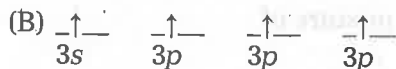
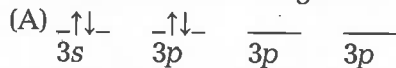


- The particulate model above is a representation of a mixture of
  - an element and a compound
  - atoms and molecules
  - four different atoms
  - two different compounds



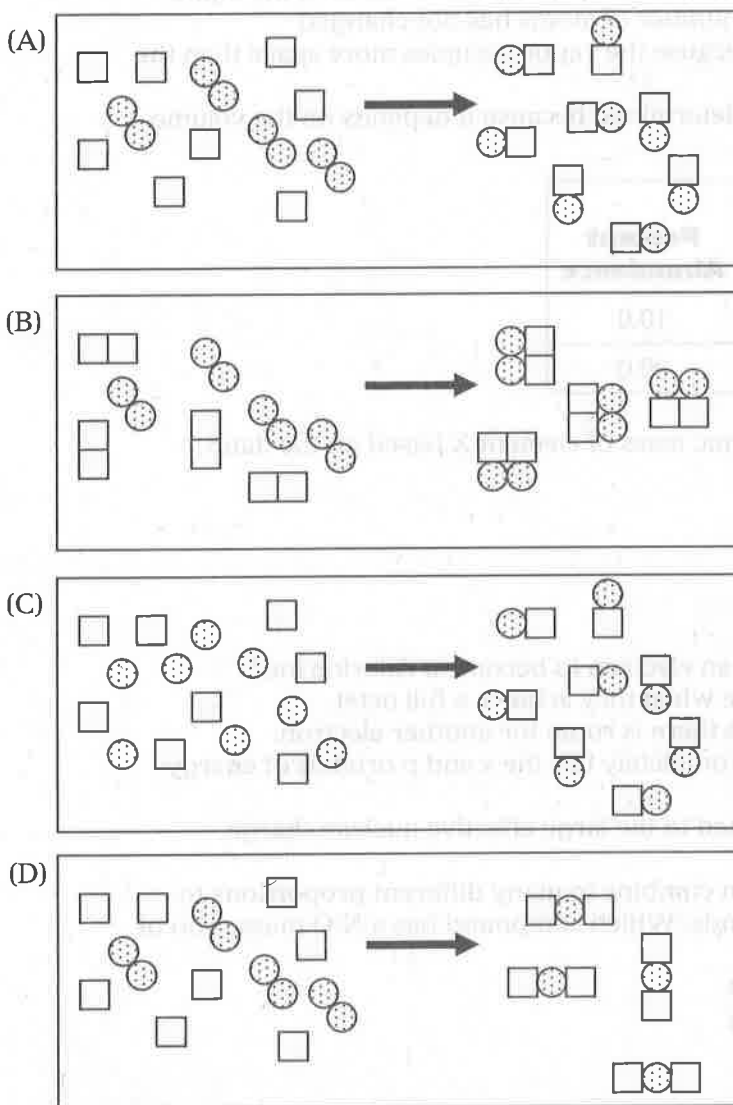
- Name the element whose mass spectrum is shown above.
  - germanium
  - selenium
  - sulfur
  - fermium
- The formation of an aluminum ion occurs when 3 electrons are removed. Which electron would require the most energy to remove?
  - 3s because it is filled before 3p
  - 3p because it is farther from the nucleus
  - 3p because it only has 1 electron in the orbital
  - 3s because it experiences higher attractive forces
- The concentration of a red-colored solution of cobalt(II) ions needs to be determined. What spectroscopic technique would be the best choice?
  - photoelectron spectroscopy because the electron arrangement is responsible for the color
  - UV-Vis spectrometry because the solution has a color
  - infrared spectrometry because the solution is red
  - mass spectrometry because knowing the mass can allow for calculation of concentration

- The correct orbital diagram for the valence electrons of silicon is





6. Glass for electronic devices needs to be more durable than typical window glass. Sodium ions on the glass surface are replaced by larger ions when the glass is dipped into a molten salt. Which type of salt would give the toughest glass?
- (A) lithium  
(B) magnesium  
(C) potassium  
(D) calcium
7. The chemical reaction  $\text{Br}_2 + 2 \text{Na} \rightarrow 2 \text{NaBr}$  is best represented by which model?



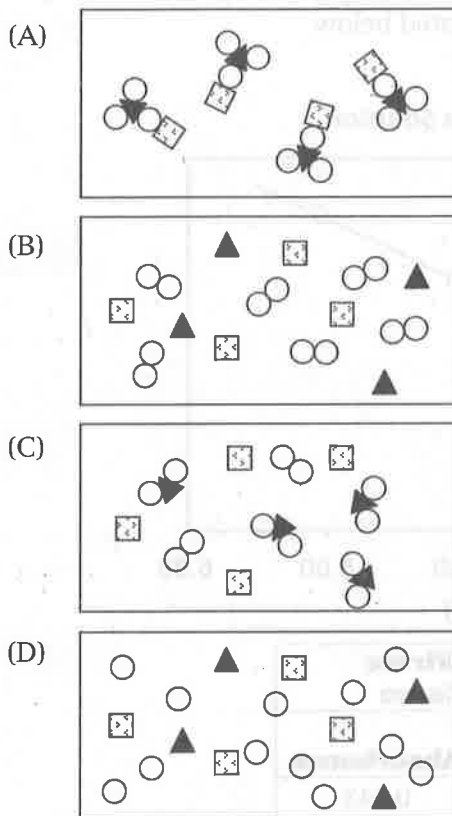
8. Sucrose,  $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ , has a molar mass of 342 g/mol. How many atoms of carbon are there in 684 g of sucrose?
- (A)  $6.02 \times 10^{23}$  atoms  
(B)  $1.45 \times 10^{25}$  atoms  
(C)  $1.20 \times 10^{24}$  atoms  
(D)  $3.01 \times 10^{23}$  atoms

9. The quantum mechanical atomic model replaced the Bohr atomic model because
- Bohr's model did not account for neutrons
  - the quantum mechanical model was based on mathematics
  - electrons act as waves as well as particles
  - Bohr's model could not predict the spectra for elements other than hydrogen
10. 1 gram of mercury is sealed in a glass tube. The tube is heated slightly and all of the mercury vaporizes. What mass of mercury remains in the tube?
- less than 1 gram because the vapor has less mass than the liquid
  - 1 gram because the number of atoms has not changed
  - more than 1 gram because the vapor occupies more space than the liquid
  - the mass cannot be determined because it depends on the volume of the tube

Isotope	Atomic Mass (amu)	Percent Abundance
$^{48}\text{X}$	48.0	10.0
$^{50}\text{X}$	50.0	90.0

11. What is the average atomic mass of element X based on the data above?
- 49.0 amu
  - 49.5 amu
  - 49.8 amu
  - 50.0 amu
12. Why does chlorine gain an electron to become a chloride ion?
- Atoms become stable when they achieve a full octet.
  - Less shielding means there is room for another electron.
  - Adding an electron completely fills the *s* and *p* orbitals of energy level 3.
  - An electron is attracted to the large effective nuclear charge.
13. Nitrogen and oxygen can combine in many different proportions to create different compounds. Which compound has a N:O mass ratio of 7:20?
- dinitrogen monoxide
  - dinitrogen pentoxide
  - nitrogen monoxide
  - nitrogen dioxide

14. Which particulate model represents copper(II) carbonate?



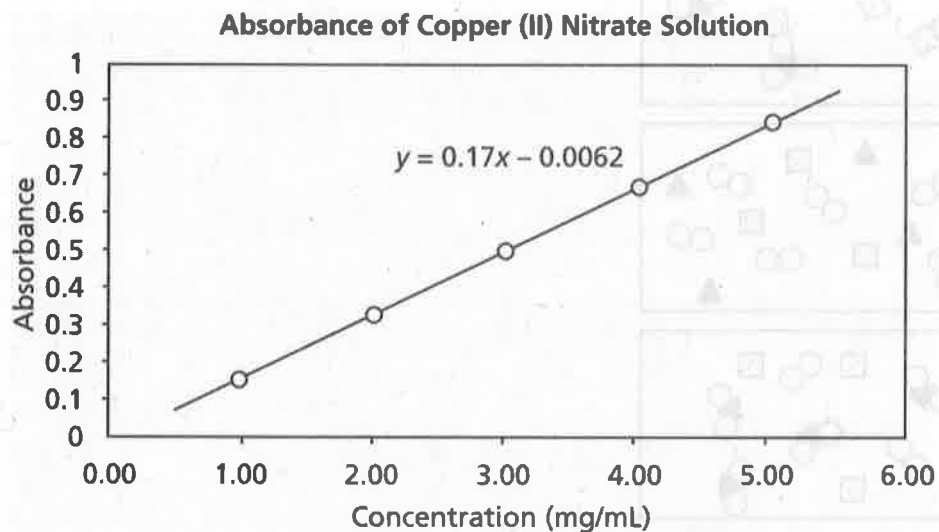
15. How many moles of sulfuric acid would completely react with 0.30 moles potassium hydroxide in a titration?

- (A) 0.15  
 (B) 0.30  
 (C) 0.60  
 (D) 1.0

### FREE-RESPONSE QUESTIONS

1. A 1956 dime is made of a silver-copper alloy and has a mass of 2.490 g. The dime was dissolved in nitric acid and a blue solution remained. The solution was quantitatively transferred to a 100-mL volumetric flask and brought to volume with water.
- (a) A 50.00-mL aliquot of the solution was reacted with excess sodium chloride solution. The precipitate, AgCl, was filtered, washed, and dried. The mass of the precipitate was 1.375 g. What is the percent silver in the dime?

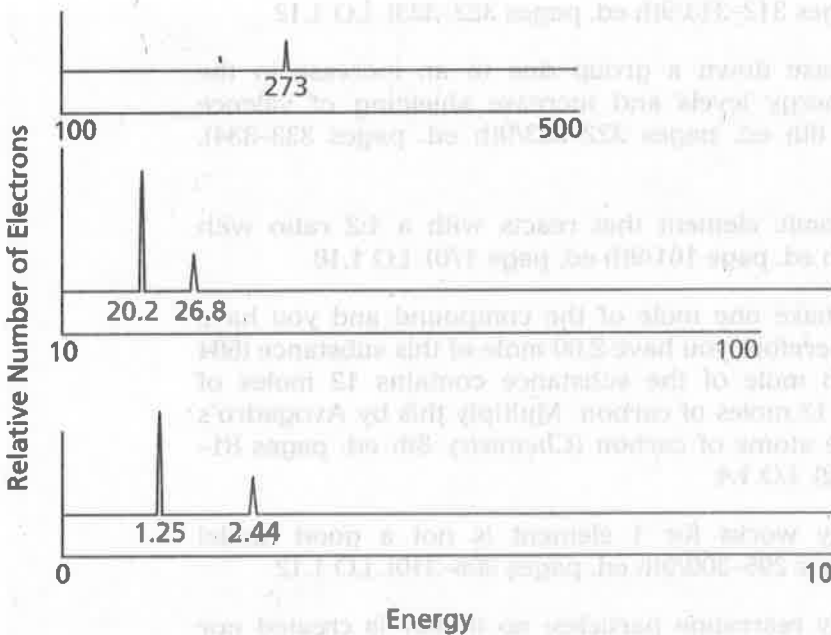
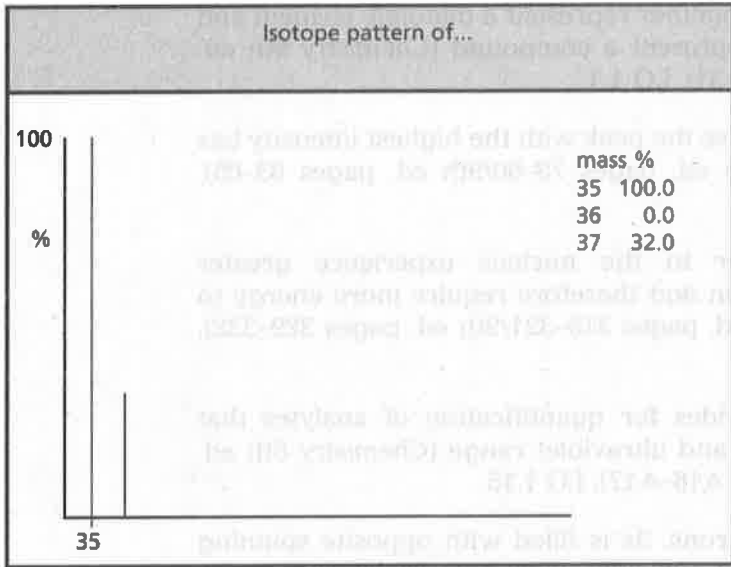
- (b) A 1.5-mL aliquot of the original sample was analyzed with a UV-Vis spectrophotometer at 740 nm and the absorbance was 0.420. Calculate the copper(II) ion concentration and the percent silver in the dime. Calibration curve data are presented below.



<b>Copper(II) Nitrate Standard Curve</b>	
<b>concentration (mg/mL)</b>	<b>Absorbance</b>
5.00	0.843
4.00	0.672
3.00	0.507
2.00	0.336
1.00	0.161

- (c) The 1956 dime is known to be 90% silver.
- (i) Which method produced more accurate results?
  - (ii) Comment on potential sources of error for the less accurate method.

2. A mysterious element was analyzed by photoelectron spectroscopy (PES) and mass spectroscopy (MS). The spectrograms are below.
- Identify the mysterious element from the spectral data.
  - Explain how the MS supports your identification.
  - Explain how the PES supports your identification.



## Answers

### MULTIPLE-CHOICE QUESTIONS

- A** Each shape in the particulate model represents a different element. The two trapezoids bonded together represent a diatomic element and the triangle-square-circles represent a compound (*Chemistry* 8th ed. pages 28–29/9th ed. pages 28–31). LO 1.1
- C** The element is sulfur because the peak with the highest intensity has a mass of 32 (*Chemistry* 8th ed. pages 78–80/9th ed. pages 83–85). LO 1.14
- D** Electrons that are closer to the nucleus experience greater Coulombic forces of attraction and therefore require more energy to be removed (*Chemistry* 8th ed. pages 318–321/9th ed. pages 329–332). LO 1.9
- B** UV-Vis spectrometry provides for quantification of analytes that absorb energy in the visible and ultraviolet range (*Chemistry* 8th ed. pages A16–A17/9th ed. pages A16–A17). LO 1.15
- D** Silicon has 4 valence electrons. 3s is filled with opposite spinning electrons and 3p has 2 filled subshells with parallel electron spins (*Chemistry* 8th ed. pages 312–313/9th ed. pages 322–323). LO 1.12
- C** Atomic radii increase down a group due to an increase in the volume of higher energy levels and increase shielding of valence electrons (*Chemistry* 8th ed. pages 322–323/9th ed. pages 333–334). LO 1.11
- A** Bromine is a diatomic element that reacts with a 1:2 ratio with sodium (*Chemistry* 8th ed. page 161/9th ed. page 170). LO 1.18
- B** It takes 342 g to make one mole of the compound and you have twice that amount. Therefore you have 2.00 mole of this substance (684 g / 342 g/mol). Each mole of the substance contains 12 moles of carbon, for a total of 12 moles of carbon. Multiply this by Avogadro's number to obtain the atoms of carbon (*Chemistry* 8th ed. pages 81–87/9th ed. pages 85–92). LO 1.4
- D** A model that only works for 1 element is not a good model (*Chemistry* 8th ed. pages 295–300/9th ed. pages 306–310). LO 1.12
- B** Phase changes only rearrange particles; no matter is created nor destroyed (*Chemistry* 8th ed. page 41/9th ed. page 44). LO 1.18
- C**  $(48.0 \text{ amu})(0.100) + (50.0 \text{ amu})(0.900) = 49.8 \text{ amu}$  (*Chemistry* 8th ed. pages 78–81/9th ed. pages 83–85). LO 1.1
- D** "A" group number = number of valence electrons (*Chemistry* 8th ed. page 316/9th ed. page 328). LO 1.9

13. **B**  $\text{N}_2\text{O}_5$ ; 28 g nitrogen and 80 g oxygen in 1 mole give a ratio of 7:20 (*Chemistry* 8th ed. page 43/9th ed. page 46). LO 1.1
14. **A** The particulate model shows only 1 kind of particle (*Chemistry* 8th ed. pages 28–29/9th ed. pages 28–31). LO 1.1
15. **A** The ratio of  $\text{H}_2\text{SO}_4$  to KOH in the balanced equation is 1 : 2 so half as many moles of acid are needed (*Chemistry* 8th ed. pages 158–159/9th ed. pages 167–168). LO 1.20

### FREE-RESPONSE QUESTIONS

1. (a)  $\text{g Ag}^+ = (1.375 \text{ g AgCl})(1 \text{ mol AgCl}/143.35 \text{ g AgCl})(1 \text{ mol Ag}^+/1 \text{ mol AgCl})$   
 $(107.2 \text{ g Ag}^+/1 \text{ mol Ag}^+)$   
 $= 1.305 \text{ g Ag}^+$   
 1.305 g  $\text{Ag}^+$  in the 50.00-mL aliquot means 2.070 g  $\text{Ag}^+$  in the 100.00-mL sample and 2.070 g Ag in the dime.
- $(2.070 \text{ g Ag}/2.490 \text{ g dime}) \times 100 = 83.13\% \text{ Ag}$   
*(Chemistry* 8th ed. pages 88–90/9th ed. pages 94–96). LO 1.19
- (b) The equation of best fit,  $y = 0.17x - 0.0062$ , shows the relationship between absorbance and concentration of  $\text{Cu}^{2+}$ .  
 $x = (0.420 + 0.0062)/0.17$   
 $x = 2.51 \text{ mg/mL}$   
 $x = 0.251 \text{ g Cu}^{2+}/100 \text{ mL solution}$ ; 0.251 g Cu in the dime  
 $(0.251 \text{ g Cu}/2.490 \text{ g dime}) \times 100 = 10.1\% \text{ Cu}$  and 89.9% Ag  
*(Chemistry* 8th ed. pages A17–A19/9th ed. pages A17–A19).  
 LO 1.16
- (c) (i) The spectrometric method was more accurate with a percent error of  $-0.1\%$  compared to  $-7.6\%$  error in the gravimetric analysis.
- (ii) Potential sources of error include loss of analyte during transfer for the reaction or loss of precipitate during filtration  
*(Chemistry* 8th ed. page A10/9th ed. page A10). LO 5.1
2. (a) The element is chlorine.
- (b) MS shows a prominent peak at 35 amu. Chlorine's average atomic mass of 35.45 amu means the majority of the isotopes are  $^{35}\text{Cl}$   
*(Chemistry* 8th ed. pages 78–80/9th ed. pages 83–85). LO 1.14
- (c) PES shows 3 occupied energy levels. The electrons closest to the nucleus have the highest ionization energy and must be 1s. The height of the top peak (energy 273 MJ) must correspond to 2 electrons in 1s. The other short peaks in the spectrum (26.8 MJ and 2.44 MJ) must also correspond to 2 electrons in s orbitals. The peak at 20.2 MJ is 3 times the height of the nearby peak meaning that peak must represent 6 electrons in 2p. The peak at 1.25 MJ is 2.5 times the height of the peak at 3s so that must represent 5 electrons in 3p. The electron configuration must be  $1s^2 2s^2 2p^6 3s^2 3p^5$ .  
 LO 1.7

