CHAPTER ONE- Matter & Measurement

AP Chemistry Chapter 1 & 2 assignments
Test & homework collection on Wednesday 8/24

1) Read chapters one and two.
2) Read the end of chapter problems for chapter 1 & 2 (not the additional exercises). These are ALL problems you need to be able to answer. Define vocabulary words you don't know and get help with problems you can't answer.
3) Complete the following problems to turn in.

Chapter 1 & 2 HOMEWORK
#1~ Pg 31- #5, 15, 16, 24, 25, 28, 29, 35, 39, 41, 45, 57, 60, 65, 70
#2~ Pg 7- #1, 9, 27, 29, 31, 32, 33, 34, 39, 41, 51, 53, 65, 69, 71, 84

Chemistry is the study of properties of materials and changes that they undergo.

1.1 The Atomic and Molecular Perspective of Chemistry

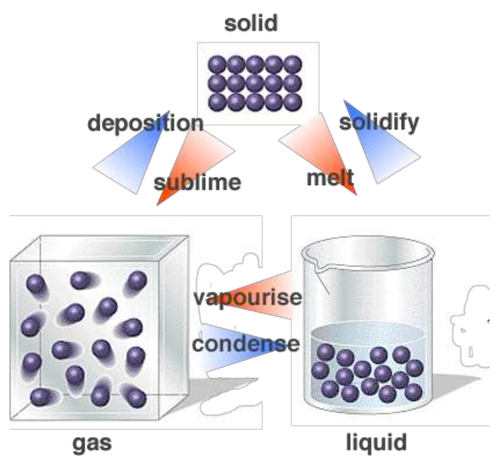
Chemistry involves the study of the properties and the behavior of matter.

Matter:
- is the physical material of the universe.
- has mass. •occupies space.
- ~100 **elements** constitute all matter.
- A **property** is any characteristic that allows us to recognize a particular type of matter and to distinguish it from other types of matter.

Elements
- are made up of unique **atoms**, the building blocks of matter.
- Element names are derived from a wide variety of sources (e.g., Latin or Greek, mythological characters, names of people or places).
- **Memorize element symbols**

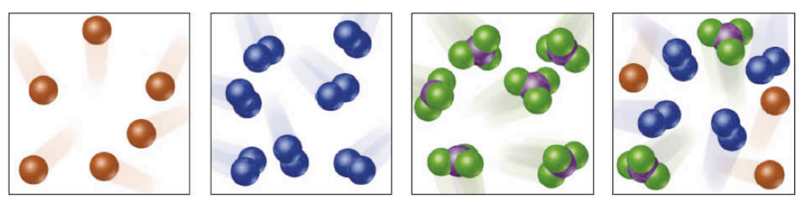
Molecules
- are combinations of atoms held together in specific shapes.
- Macroscopic (observable) properties of matter relate to submicroscopic realms of atoms.
- Properties relate to composition (types of atoms present) and structure (arrangement of atoms) present.

1.2 Classifications of Matter-- Matter is classified by *state* (solid, liquid or gas) or by *composition* (element, compound or mixture).



States of Matter
- Solids, liquids and gases are the three forms of matter called the **states of matter**.
- Properties described on the macroscopic level:
  - •**gas** (vapor): no fixed volume or shape, conforms to shape of container, compressible.
  - •**liquid**: volume independent of container, no fixed shape, incompressible.
  - •**solid**: volume and shape independent of container, rigid, incompressible.
- Properties described on the molecular level:
  - •**gas**: molecules far apart, move at high speeds, collide often.
  - •**liquid**: molecules closer than gas, move rapidly but can slide over each other.
  - •**solid**: molecules packed closely in definite arrangements.

Pure substances:
- are matter with fixed compositions and distinct proportions.
- are **elements** (cannot be decomposed into simpler substances, i.e. only one kind of atom) or **compounds** (consist of two or more elements).



(a) Atoms of an element (b) Molecules of an element (c) Molecules of a compound (d) Mixture of elements and a compound

Elements
- There are 118 known elements.
- They vary in abundance.
- Each is given a unique name and is abbreviated by a chemical *symbol*.
- they are organized in the periodic table.
- Each has a one- or two-letter symbol derived from its name.

Compounds
- **Compounds** are combinations of elements. (The compound H_2O is a combination of the elements H and O.)
- The opposite of compound formation is decomposition.
- Compounds have different properties than their component elements (e.g., water is liquid, but hydrogen and oxygen are both gases at the same temperature and pressure).
- **Law of Constant (Definite) Proportions** (Proust): A compound always consists of the same combination of elements (water is always 11% H and 89% O).

Mixtures
- A **mixture** is a combination of two or more pure substances.
- Each substance retains its own identity; each substance is a *component* of the mixture.
- Mixtures have variable composition.
- Heterogeneous mixtures do not have uniform composition, properties, and appearance, e.g., sand.
- Homogeneous mixtures are uniform throughout, e.g., air; they are **solutions**.

1.3 Properties of Matter

Each substance has a unique set of physical and chemical properties.
- **Physical properties** are measured without changing the substance (e.g., color, density, odor, melting point, etc.).
- **Chemical properties** describe how substances react or change to form different substances

Properties may be categorized as intensive or extensive.
- **Intensive properties** do not depend on the amount of substance present (e.g., temperature, melting point etc.).
- **Extensive properties** depend on the quantity of substance present (e.g., mass, volume etc.).
- Intensive properties give an idea of the composition of a substance •Extensive properties give an indication of the quantity of substance.

Physical and Chemical Changes

- **Physical change**: substance changes physical appearance without altering its identity (e.g., **changes of state**).
- **Chemical change** (or **chemical reaction**): substance transforms into a chemically different substance (i.e. identity changes, e.g., decomposition of water, explosion of nitrogen triiodide).

Separation of Mixtures Key: separation techniques exploit differences in properties of the *components*.

- Filtration: remove solid from liquid.
- Distillation: boil off one or more components of the mixture.
- Chromatography: exploit solubility of components.

The Scientific Method provides guidelines for the practice of science.

- Collect data (observe, experiment, etc.).
- Look for patterns, try to explain them, and develop a **hypothesis** *or tentative explanation*.
- Test hypothesis, then refine it.
- Bring all information into a **scientific law** (concise statement or equation that summarizes tested hypotheses).
- Bring hypotheses and laws together into a theory. A **theory** should explain general principles.

PRACTICE EXERCISE

(a) Convert 2700ns to seconds.

(b) Express the measurement 6.0×10^3 m using a prefix to replace the power of ten.

(c) Use exponential notation to express 3.76 mg in grams.

1.4 Units of Measurement

- Many properties of matter are quantitative, i.e., associated with numbers.
- A measured quantity must have BOTH a number and a unit.
- The units most often used for scientific measurement are those of the **metric system**.

SI Units

- 1960: All scientific units use **S**ystème **I**nternational d'Unités (**SI** Units).
- There are seven base units.
- Smaller and larger units are obtained by decimal fractions or multiples of the base units.

Length and Mass

- SI base unit of length = meter (1 m = 1.0936 yards).
- SI base unit of mass (not weight) = kilogram (1 kg = 2.2 pounds).
  - **Mass** is a measure of the amount of material in an object.

Temperature is the measure of the hotness or coldness of an object.

- Scientific studies use Celsius and Kelvin scales.
  - **Celsius scale**: water freezes at 0ºC and boils at 100ºC (sea level).
  - **Kelvin scale** (SI Unit):
    - Water freezes at 273.15 K and boils at 373.15 K (sea level).
    - is based on properties of gases.
    - Zero Kelvin is the lowest possible temperature (absolute zero is -273.15°C).
  - **Fahrenheit (not used in science):**
    - Water freezes at 32ºF and boils at 212ºF (sea level).

$$°F = \frac{9}{5}°C + 32$$

$$°C = \frac{5}{9}(°F - 32)$$

$$°C = K - 273.15$$

$$K = °C + 273.15$$

Derived SI Units

- These are formed from the seven base units.
- Velocity is distance traveled per unit time, so units of velocity are units of distance (m) divided by units of time (s): m/s.

Volume

- Units of volume = (units of length)^3 = m^3.
- This unit is unrealistically large, so we use more reasonable units:
  - cm^3 [also known as mL (milliliter) or cc (cubic centimeters)]
  - dm^3 (also known as liters, L). **Important: the liter is not an SI unit.**

Density

- Is used to characterize substances.
- **Density** is defined as mass divided by volume.
  - Units: g/cm^3 or g/mL (for solids and liquids); g/L (often used for gases).
  - density of water at 25°C is exactly 1g/mL. The gram is defined as the mass of 1.00 mL of pure water at 25ºC.

SAMPLE EXERCISE 1.4

(a) Calculate the density of mercury if 1.00×10^2 g occupies a volume of 7.36 cm^3.

(b) Calculate the volume of 65.0 g of the liquid methanol (wood alcohol) if its density is 0.791 g/mL.

(c) What is the mass in grams of a cube of gold (density = 19.32 g/ cm^3) if the length of the cube is 2.00 cm?

1.5 Uncertainty in Measurement

- There are two types of numbers:
  - *exact numbers* (known as counting or defined).
  - *inexact numbers* (derived from measurement).
- All measurements have some degree of uncertainty or *error* associated with them.

Precision and Accuracy

- **Precision**: how well measured quantities agree with each other.
- **Accuracy**: how well measured quantities agree with the "true value."

Significant Figures

•In a measurement the exactness of the measurement is reflected in the number of significant figures.

•Guidelines for determining the number of significant figures in a measured quantity are:

•The number of significant figures is the number of digits known with certainty plus one uncertain digit. (2.2405 g means we are sure the mass is 2.240 g but we are uncertain about the nearest 0.0001 g.)

•Final calculations are only as significant as the least significant measurement.

When in doubt, report 3 sig figs.

Rules:
1. Nonzero numbers and zeros between nonzero numbers are always significant.
2. Zeros before the first nonzero digit are not significant. (Example: 0.0003 has one significant figure.)
3. Zeros at the end of the number after a decimal point are significant.
4. Zeros at the end of a number before a decimal point are ambiguous (ex. 10,300 g). Exponential notation eliminates this ambiguity.

SAMPLE EXERCISE 1.6
How many significant figures are in each of the following numbers (assume that each number is a measured quantity):
(a) 4.003
(b) 6.023×10^{23}
(c) 5000
(d) 0.0760

Significant Figures in Calculations

•Multiplication and division:

•Report to the least number of significant figures (e.g., 6.221 cm x 5.2 cm = 32 cm^2).

•Addition and subtraction:

•Report to the least number of decimal places (e.g., 20.4 g – 1.322 g = 19.1 g).

•In multiple step calculations always retain an extra significant figure until the end to prevent rounding errors.

SAMPLE EXERCISE 1.7
The width, length, and height of a small box are 15.5 cm, 27.3 cm, and 5.4 cm, respectively. Calculate the volume of the box, using the correct number of significant figures in your answer.

SAMPLE EXERCISE 1.8 A gas at 25°C fills a container whose volume is 1.05×10^3 cm^3. The container plus gas have a mass of 837.6 g. The container, when emptied of all gas, has a mass of 836.2 g. What is the density of the gas at 25°C?

1.6 Dimensional Analysis

Dimensional analysis is a method of calculation utilizing a knowledge of units.

•Conversion factors are used to manipulate units.

•desired unit = given unit x (conversion factor)

•The **conversion factors** are simple ratios.

•conversion factor = (desired unit) / (given unit)

•These are fractions whose numerator and denominator are the same quantity in different units.

•Multiplication by a conversion factor is equivalent to multiplying by a factor of one.

Using Two or More Conversion Factors- We often need to use more than one conversion factor in order to complete a problem.

- When identical units are found in the numerator and denominator of a conversion chain, they will cancel. The final answer MUST have the correct units.

Conversions Involving Volume

•Suppose that we wish to know the mass in grams of 2.00 cubic inches of gold given that the density of the gold is 19.3 g/cm^3.

•We could do this conversion with the following conversion factors:

2.54 cm = 1 inch and 1 cm^3 = 19.3 g gold

•The calculation would involve both of these factors:

$(2.00\ in.^3)\ (2.54\ cm / in.)^3\ (19.3\ g\ gold / 1\ cm^3) = 633\ g\ gold$

•Note that the calculation will NOT be correct unless the centimeter to inch conversion factor is cubed!! Both the units AND the number must be cubed.

Summary of Dimensional Analysis

•In dimensional analysis always ask three questions:
1. What data are we given?
2. What quantity do we need?
3. What conversion factors are available to take us from what we are given to what we need?

SAMPLE EXERCISE 1.9
If a woman has a mass of 115 lb, what is her mass in grams? (Use the relationships between units given on the back inside cover of the text.)

SAMPLE EXERCISE 1.10
The average speed of a nitrogen molecule in air at 25°C is 515 m/s. Convert this speed to miles per hour.

SAMPLE EXERCISE 1.11
Earth's oceans contain approximately 1.36×10^9 km^3 of water. Calculate the volume in liters.

CHAPTER TWO- Atoms, Molecules, & Ions

2.1 The Atomic Theory of Matter

•Democritus (460–370 BC): All matter can be divided into indivisible *atomos*.

•Dalton: proposed atomic theory with the following postulates:

•Elements are composed of atoms.

•All atoms of an element are identical.

•In chemical reactions atoms are not changed into different types of atoms. Atoms are neither created nor destroyed.

•Compounds are formed when atoms of elements combine.

•**Atoms** are the building blocks of matter.

•*Law of constant composition*: The relative kinds and numbers of atoms are constant for a given compound.

•*Law of conservation of mass (matter)*: For chemical reactions, the total mass before the reaction is equal to the total mass after the reaction.

•Conservation means something can neither be created nor destroyed. Here, it applies to matter (mass). Later we will apply it to energy (Chapter 5).

•*Law of multiple proportions*: If two elements, A and B, combine to form more than one compound, then the mass of B, which combines with the mass of A, is a ratio of small whole numbers.

•Dalton's theory *predicted* the law of multiple proportions.

2.2 The Discovery of Atomic Structure

•By 1850 scientists knew that atoms consisted of charged particles.

•**Subatomic particles** are those particles that make up the atom.

•Recall the law of electrostatic attraction**:** like charges repel and opposite charges attract.

Cathode Rays and Electrons

•Cathode rays were first discovered in the mid-1800s from studies of electrical discharge through partially evacuated tubes (cathode-ray tubes or CRTs).

•Computer terminals were once popularly referred to as CRTs (cathode-ray tubes).

•Cathode rays = radiation produced when high voltage is applied across the tube.

•The voltage causes negative particles to move from the negative electrode to the positive electrode.

•The path of the electrons can be altered by the presence of a magnetic field.

•Consider cathode rays leaving the positive electrode through a small hole.

•If they interact with a magnetic field perpendicular to an applied electric field, then the cathode rays can be deflected by different amounts.

•The amount of deflection of the cathode rays depends on the applied magnetic and electric fields.

•In turn, the amount of deflection also depends on the charge-to-mass ratio of the electron.

•**In 1897 Thomson determined the charge-to-mass ratio of an electron in his cathode ray experiment.**

•Charge-to-mass ratio: 1.76×10^8 C/g. C is a symbol for coulomb (the SI unit for electric charge)

Millikan Oil-Drop Experiment

•Goal: find the charge on the electron to determine its mass.

•Oil drops are sprayed above a positively charged plate containing a small hole.

•As the oil drops fall through the hole they acquire a negative charge.

•Gravity forces the drops downward. The applied electric field forces the drops upward.

•When a drop is perfectly balanced, then the weight of the drop is equal to the electrostatic force of attraction between the drop and the positive plate.

•Millikan determined the charges on the oil drops to be multiples of 1.60×10^{-19} C.

•He concluded the charge on the electron must be 1.60×10^{-19} C.

•Knowing the charge-to-mass ratio of the electron, we can calculate the mass of the electron:

$$\text{Mass} = \frac{1.60\times10^{-19}\,\text{C}}{1.76\times10^{8}\,\text{C/g}} = 9.10\times10^{-28}\,\text{g}$$

Radioactivity is the spontaneous emission of radiation.

•Consider the following experiment:

•A radioactive substance is placed in a lead shield containing a small hole so that a beam of radiation is emitted from the shield.

•The radiation is passed between two electrically charged plates and detected.

•Three spots are observed on the detector:

1. a spot deflected in the direction of the positive plate,
2. a spot that is not affected by the electric field, and
3. a spot deflected in the direction of the negative plate.

•A large deflection towards the positive plate corresponds to radiation that is negatively charged and of low mass. This is called β-radiation (consists of electrons).

•No deflection corresponds to neutral radiation. This is called γ-radiation (similar to X-rays).

•A small deflection toward the negatively charged plate corresponds to high mass, positively charged radiation. This is called α-radiation (positively charged core of a helium atom)

•X-rays and γ radiation are electromagnetic radiation, whereas α- and β-radiation are streams of particles

The Nuclear Atom

•The plum pudding model is an early picture of the atom.

•The Thomson model pictures the atom as a sphere with small electrons embedded in a positively charged mass.

•Rutherford carried out the following "gold foil" experiment:

•α-particles were shot through a piece of gold foil.

•Both the gold nucleus and the α-particle were positively charged, so they repelled each other.

•Most of the α-particles went straight through the foil without deflection.

•If the Thomson model of the atom was correct, then Rutherford's result was impossible.

•Rutherford modified Thomson's model as follows:

•The positive charge must be located at the center with a diffuse negative charge surrounding it.

•To account for the small number of large deflections of the α-particles, the center or **nucleus** of the atom must consist of a dense positive charge.

2.3 The Modern View of Atomic Structure

The atom consists of positive, negative, and neutral entities (**protons**, **electrons** and **neutrons**).

•Protons & neutrons are located in the nucleus, which is small. Most of the atom's mass is due to the nucleus.

•Electrons are located outside of the nucleus. Most of the volume of the atom is due to electrons.

•The quantity 1.602×10^{-19} C is called the **electronic charge**. The charge on an electron is -1.602×10^{-19} C; the charge on a proton is $+1.602 \times 10^{-19}$ C; neutrons are uncharged.

•Atoms have an equal number of protons and electrons thus they have no net electrical charge.

Masses of subatomic particles are so small that we define the **atomic mass unit**, amu.

•1 amu = 1.66054×10^{-24} g.

•The mass of a proton is 1.0073 amu, a neutron is 1.0087 amu, and an electron is 5.486×10^{-4} amu.

•The **angstrom** is a convenient non SI unit of length used to denote atomic dimensions.
•Since most atoms have radii around 1 x 10^{-10} m, we define 1 Å = 1 x 10^{-10} m.

Atomic Numbers, Mass Numbers, And Isotopes

•**Atomic number** (Z) = number of protons in the nucleus.
•**Mass number** (A) = total number of nucleons in the nucleus (i.e., protons and neutrons).
•By convention, for element X, we write $\longrightarrow \; ^{A}_{Z}\text{X}$
•Thus, isotopes have the same Z but different A.

• All atoms of a specific element have the same number of protons.
•**Isotopes** of a specific element differ in the number of neutrons. An atom of a specific isotope is called a nuclide .
Nuclides of hydrogen include: ^{1}H = hydrogen (protium), ^{2}H = deuterium, ^{3}H = tritium; tritium is radioactive.

SAMPLE EXERCISE 2.2 How many protons, neutrons, and electrons are in
(a) an atom of ^{197}Au

(b) an atom of strontium-90?

2.4 Atomic Weights
The Atomic Mass Scale
•*Atomic mass units* (amu) are convenient units to use when dealing with extremely small masses of individual atoms.
•1 amu = 1.66054 x 10^{-24} g and 1 g = 6.02214 x 10^{23} amu
•By definition, the mass of ^{12}C is exactly 12 amu.

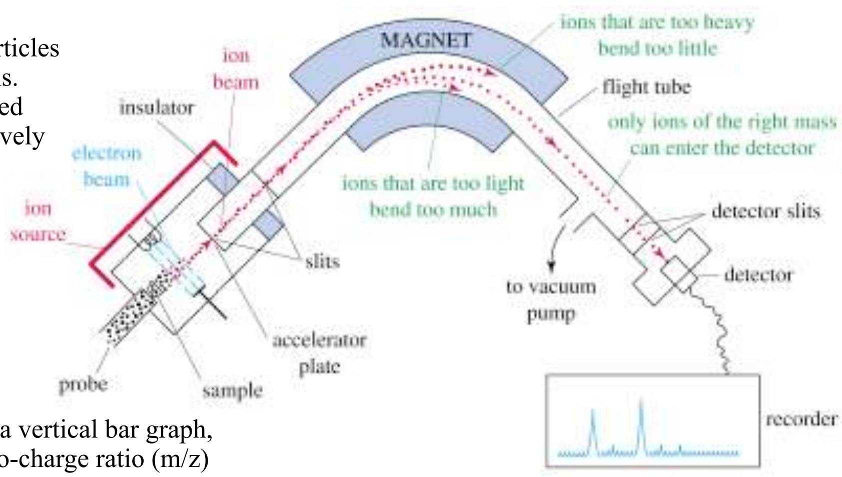
Average Atomic Masses-We average the masses of isotopes to give average atomic masses.

- Naturally occurring C consists of 98.93% ^{12}C (12 amu) and 1.07% ^{13}C (13.00335 amu).
- The average mass of C is:
  (0.9893)(12 amu) + (0.0107)(13.00335 amu) = 12.01 amu.
- **Atomic weight** is also known as average atomic mass

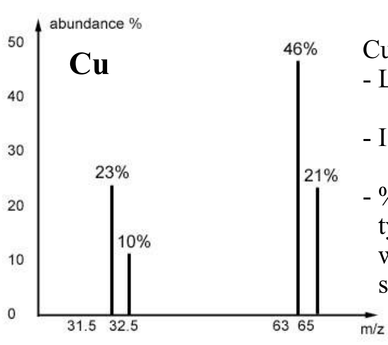
SAMPLE EXERCISE 2.4
Naturally occurring chlorine is 75.78% ^{35}Cl (34.969 amu) and 24.22% ^{37}Cl (36.966 amu). Calculate the average atomic mass (that is, the atomic weight) of chlorine.

The Mass Spectrometer is an instrument that allows for direct and accurate determination of atomic (and molecular) weights.
THREE DISTINCT PARTS

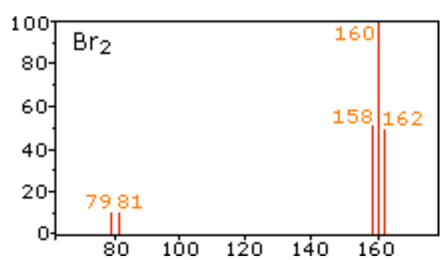
- ION SOURCE- Gaseous particles (atoms or molecules) are bombarded by high-energy electrons. This breaks the particles into all possible combinations of cations and anions/electrons. Repeller plates and accelerator plates direct positively charged particles into the mass analyzer while also preventing negatively charged particles from entering the analyzer.
- MASS ANALYZER- Positively charged particles of various sizes and charges are focused in a narrow beam which passes through a magnetic field. The path of charged particles bends as they pass through the magnetic field. **Higher charges bend more. Lighter particles bend more.** Only ions of the right mass can enter the detector, but the strength of the magnetic field can be changed, such that all particles can be detected under an appropriate field strength.
- DETECTOR- A mass spectrum will usually be presented as a vertical bar graph, in which each bar represents an ion having a specific mass-to-charge ratio (m/z) and the length of the bar indicates the relative abundance of the ion. The most intense ion is assigned an abundance of 100, and it is referred to as the **base peak**. Most of the ions formed in a mass spectrometer have a single charge, so the m/z value is equivalent to mass itself.



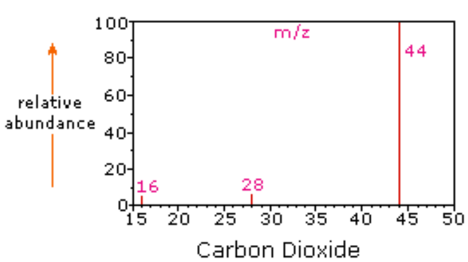
Mass Spectrometer

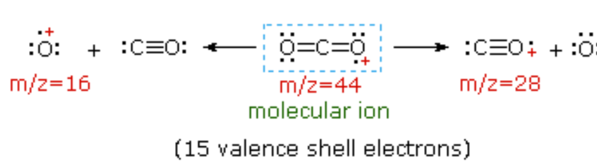


Cu Mass Spectrum
- Lower m/z values are for 2$^+$ charged particles
- Isotope masses of 63amu and 65amu are detected
- % values are NOT the typical **relative abundance,** when the highest value is scaled to 100.



Br_2 Mass Spectrum
- Only shows 1$^+$ cations, typical
- Detects two isotopes of equal abundancies and all mass combos of the diatomic element.
- It is twice as likely to have a diatomic molecule with one of each isotope than to have a molecule with two nuclides of the same mass.





Mass Spectra for Molecules are more complex, as the number of possible positive fragments increases with the complexity of the molecule.

Each substance has a unique mass spectrum, so MS is used to identify compounds and detect impurities.

2.5 The Periodic Table

The **periodic table** is used to organize the elements in a meaningful way.

- As a consequence of this organization, there are periodic properties associated with the periodic table.

Rows in the periodic table are called periods.

Columns in the periodic table are called **groups**.

- Several numbering conventions are used (i.e., groups may be numbered from 1 to 18, or from 1A to 8A and 1B to 8B).
- Some of the groups in the periodic table are given special names.
    - These names indicate the similarities between group members.
    - Examples: •Group 1A: alkali metals •Group 2A: alkaline earth metals •Group 7A: halogens •Group 8A: noble gases

Metallic elements, or **metals**, are located on the left side of the periodic table (most elements are metals).

- Metals tend to be malleable, ductile, and lustrous and are good thermal and electrical conductors.

Nonmetallic elements, or **nonmetals**, are located in the top right-hand side of the periodic table.

- Nonmetals tend to be brittle as solids, dull in appearance, and do not conduct heat or electricity well.

Elements with properties similar to both metals and nonmetals are called **metalloids** and are located at the interface between the metals and nonmetals. These include the elements B, Si, Ge, As, Sb and Te.

2.6 Molecules and Molecular Compounds A molecule consists of two or more atoms bound tightly together.

Molecules and Chemical Formulas

- Each molecule has a **chemical formula** that indicates
    1. which atoms are found in the molecule, and
    2. in what proportion they are found.
- A molecule made up of two atoms is called a **diatomic molecule**.
- Different forms of an element, which have different chemical formulas, are known as allotropes. Allotropes differ in their chemical and physical properties.
- Compounds composed of molecules are **molecular compounds**.
    - These contain at least two types of atoms.
    - Most molecular substances contain only nonmetals.

SAMPLE EXERCISE 2.6
Give the empirical formula for the following compounds:
(a) $C_6H_{12}O_6$
(b) N_2O
(c) B_2H_6

Molecular and Empirical Formulas

- **Molecular formulas**
    - These formulas give the actual numbers and types of atoms in a molecule.
    - Examples: H_2O, CO_2, CO, CH_4, H_2O_2, O_2, O_3, and C_2H_4.
- **Empirical formulas**
    - These formulas give the relative numbers and types of atoms in a molecule (the lowest whole-number ratio of atoms in a molecule).
    - Examples: H_2O, CO_2, CO, CH_4, HO, CH_2.

Picturing Molecules Molecules occupy three-dimensional space. However, we often represent them in two dimensions.

- The **structural formula** gives the connectivity between individual atoms in the molecule.
- The structural formula can show the three-dimensional shape of the molecule.
- If the structural formula does show the shape of the molecule then either a perspective drawing, a ball-and-stick model, or a space-filling model is used.
    - *Perspective drawings* use dashed lines and wedges to represent bonds receding and emerging from the plane of the paper.
    - *Ball-and-stick models* show atoms as spheres and the bonds as sticks. The angles in the ball-and-stick model are accurate.
    - *Space-filling models* give an accurate representation of the 3-D shape of the molecule.

2.7 Ions and Ionic Compounds

- If electrons are added to or removed from a neutral atom, an **ion** is formed.
- When an atom or molecule loses electrons it becomes positively charged. Positively charged ions are called **cations**.
- When an atom or molecule gains electrons it becomes negatively charged. Negatively charged ions are called **anions**.
- In general, metal atoms tend to lose electrons and nonmetal atoms tend to gain electrons.
- When molecules lose electrons, **polyatomic ions** are formed (e.g. SO_4^{2-}, NO_3^-).

SAMPLE EXERCISE 2.6 Give the chemical symbol, including mass number, for the ion of sulfur that has 16 neutrons and 18 electrons.

How many protons and electrons does the Se^{2-} ion possess?

Predicting Ionic Charges

- An atom or molecule can lose more than one electron.
- Many atoms gain or lose enough electrons to have the same number of electrons as the nearest noble gas (group 8A).
- The number of electrons an atom loses is related to its position on the periodic table.

SAMPLE EXERCISE 2.8
Predict the charge expected for the most stable ion of the following:

barium	oxygen
aluminum	fluorine

Ionic Compounds

- A great deal of chemistry involves the transfer of electrons between species.

Example: •To form NaCl, the neutral sodium atom, Na, must lose an electron to become a cation: Na^+.
•The electron cannot be lost entirely, so it is transferred to a chlorine atom, Cl, which then becomes an anion: Cl^-.
•The Na^+ and Cl^- ions are attracted to form an ionic NaCl lattice, which crystallizes.

- NaCl is an example of an **ionic compound** made of positively charged cations and negatively charged anions.
    - Important: note that there are no easily identified NaCl molecules in the ionic lattice. Therefore, we cannot use molecular formulas to describe ionic substances.
- Ionic compounds are combinations of metals and nonmetals, and molecular compounds are composed of nonmetals only.

SAMPLE EXERCISE 2.9 Which of the following compounds are ionic: N_2O, Na_2O, $CaCl_2$, SF_4, CBr_4, FeS, P_4O_6, PbF_2?

•Writing empirical formulas for ionic compounds:
- •You need to know the ions of which it is composed.
- •The formula must reflect the electrical neutrality of the compound.
  - •You must combine cations and anions in a ratio so that the total positive charge is equal to the total negative charge.
    - •Example: Consider the formation of Mg_3N_2:
      - •Mg loses two electrons to become Mg^{2+}
      - •Nitrogen gains three electrons to become N^{3-}.
      - •For a neutral species, the number of electrons lost and gained must be equal.
      - •However, Mg can only lose electrons in twos and N can only accept electrons in threes.
      - •Therefore, Mg needs to lose six electrons (2x3) and N gains those six electrons (3x2). That is, 3Mg atoms need to form $3Mg^{2+}$ ions (total 3x2 positive charges) and 2N atoms need to form $2N^{3-}$ ions (total 2x3 negative charges).
      - •Therefore, the formula is Mg_3N_2.

Chemistry and Life: Elements Required by Living Organisms
- •Of the 116 elements known, only about 29 are required for life.
- •Water accounts for at least 70% of the mass of most cells.
- •Carbon is the most common element in the solid components of cells.
- •The most important elements for life are H, C, N, O, P and S. The next most important ions are Na^+, Mg^{2+}, K^+, Ca^{2+}, and Cl^- .
- •The other required 18 elements are only needed in trace amounts; they are *trace elements*.

2.8 Naming Inorganic Compounds
•**Chemical nomenclature** is the naming of substances.
•Common names are traditional names for substances (e.g., water, ammonia).
•Systematic names are based on a systematic set of rules.
- •Divided into organic compounds (those containing C, usually in combination with H, O, N, or S) and inorganic compounds (all others).

Names and Formulas of Ionic Compounds
Ions formed from a single atom are called *monoatomic ions*.

1. Positive Ions (Cations)
- •Cations formed from a metal have the same name as the metal.
  - •Example: Na^+ = sodium ion.
- •Many transition metals exhibit variable charge.
  - •If the metal can form more than one cation, then the charge is indicated in parentheses in the name.
    - •Examples: Cu^+ = copper(I) ion; Cu^{2+} = copper(II) ion.
- •Cations formed from nonmetals end in **-ium**.
  - •Examples: NH_4^+ = ammonium ion; H_3O^+ = hydronium ion.

2. Negative Ions (Anions)
- •Monatomic anions (with only one atom) use the ending **-ide**.
  - •Example: Cl^- is the chloride ion.
- •Some polyatomic anions also use the -ide ending:
  - •Examples: hydroxide, cyanide, and peroxide ions.
- •Polyatomic anions (with many atoms) containing oxygen are called **oxyanions**.
  - •Their names end in -**ate** or -**ite**. (The one with more oxygen is called **-ate**.)
  - •Examples: NO_3^- is nitr**ate**; NO_2^- is nitr**ite**.
- •Polyatomic anions containing oxygen with more than two members in the series are named as follows (in order of **decreasing oxygen**):

• **per-**….**-ate**	example:	ClO_4^-	**per**chlor**ate**
• **-ate**		ClO_3^-	chlor**ate**
• **-ite**		ClO_2^-	chlor**ite**
• **hypo-**….**-ite**		ClO^-	**hypo**chlor**ite**

- •Polyatomic anions containing oxygen with additional hydrogens are named by adding hydrogen or bi- (one H), dihydrogen (two H) etc., to the name as follows:
  - • CO_3^{2-} is the carbon**ate** anion.
  - • HCO_3^- is the hydrogen carbonate (or **bi**carbonate) anion.
  - • PO_4^{3-} is the phosphate ion.
  - • $H_2PO_4^-$ is the **dihydrogen** phosphate anion.

3. Ionic Compounds
- •These are named by the cation then the anion.
  - •Example: $BaBr_2$ = barium bromide.

Names and Formulas of Acids
- •Acids are substances that yield hydrogen ions when dissolved in water (Arrhenius definition).
  - •The names of acids are related to the names of anions:

•**-ide** becomes **hydro-**….**-ic** acid;	example:	HCl	**hydro**chlor**ic** acid
•**-ate** becomes **-ic** acid;		$HClO_4$	perchlor**ic** acid
•**-ite** becomes -**ous** acid.		HClO	hypochlor**ous** acid

Names and Formulas of Binary Molecular Compounds
•Binary molecular compounds have two elements.
•The most metallic element (i.e., the one to the farthest left on the periodic table) is usually written first. The exception is NH_3.
•If both elements are in the same group, the lower one is written first.
•Greek prefixes are used to indicate the number of atoms (e.g., mono, di, tri).
•The prefix mono is never used with the first element (i.e., carbon monoxide, CO).
•Examples:
• Cl_2O is **di**chlorine *mono*xide.
• N_2O_4 is **di**nitrogen *tetr*oxide.
• NF_3 is nitrogen *tri*fluoride.
• P_4S_{10} is **tetra**phosphorus *deca*sulfide.

2.9 Some Simple Organic Compounds
Organic chemistry is the study of carbon-containing compounds.
•*Organic compounds* are those that contain carbon and hydrogen, often in combination with other elements.
Alkanes
•Compounds containing only carbon and hydrogen are called **hydrocarbons**.
•In **alkanes** each carbon atom is bonded to four other atoms.
•The names of alkanes end in *-ane*. •Examples: methane, ethane, propane, butane.
Some Derivatives of Alkanes
When *functional groups*, specific groups of atoms, are used to replace hydrogen atoms on alkanes, new classes of organic compounds are obtained.
•**Alcohols** are obtained by replacing a hydrogen atom of an alkane with an –OH group.
•Alcohol names derive from the name of the alkane and have an *-ol* ending.
•Examples: methane becomes methanol; ethane becomes ethanol.
Carbon atoms often form compounds with long chains of carbon atoms.
•Properties of alkanes and derivatives change with changes in chain length.
•*Polyethylene*, a material used to make plastic products, is an alkane with thousands of carbons.
•This is an example of a *polymer*.
Carbon may form *multiple bonds* to itself or other atoms.

Chapter One Practice Test

1) The density of silver is $10.5\,g/cm^3$. A piece of silver that occupies a volume of $23.6\,cm^3$ would have a mass of _____g.
A) 248 B) 0.445 C) 2.25
D) 112 E) 23.6

2) __________ significant figures should be retained in the result of the following calculation.
$$\frac{(11.13 - 2.6) \times 10^4}{(103.05 + 16.9) \times 10^{-6}}$$
A) 1 B) 2 C) 3
D) 4 E) 5

3) The density of mercury is $13.6\,g/cm^3$. The density of mercury is __________kg/m^3.
A) 1.36×10^{-2} B) 1.36×10^{4}
C) 1.36×10^{8} D) 1.36×10^{-5}
E) 1.36×10^{-4}

4) There are __________ ng in a pg.
A) 0.001 B) 1000
C) 0.01 D) 100 E) 10

5) If matter is uniform throughout and cannot be separated into other substances by physical means, it is _________.
A) a compound
B) either an element or a compound
C) a homogeneous mixture
D) a heterogeneous mixture
E) an element

6) Of the following, only __________ is a chemical reaction.
A) melting of lead
B) dissolving sugar in water
C) tarnishing of silver
D) crushing of stone
E) dropping a penny into a glass of water

7) Which one of the following is an intensive property?
A) mass B) temperature
C) heat content D) volume
E) amount

8) Accuracy refers to __________.
A) how close a measured number is to zero
B) how close a measured number is to the calculated value
C) how close a measured number is to other measured numbers
D) how close a measured number is to the true value
E) how close a measured number is to infinity

9) In which one of the following numbers are all of the zeros significant?
A) 100.090090
B) 0.143290
C) 0.05843
D) 1000
E) 00.0030020

10) A common English set of units for expressing velocity is miles/hour. The SI unit for velocity is __________?
A) km/hr B) km/s
C) m/hr D) m/s E) g/m

11) A temperature of __________ K is the same as 63°F.
A) 17 B) 276
C) 290 D) 29 E) 336

12) A combination of sand, salt, and water is an example of a ___.
A) homogeneous mixture
B) heterogeneous mixture
C) compound
D) pure substance
E) solid

13) Which states of matter are significantly compressible?
A) gases only
B) liquids only
C) solids only
D) liquids and gases
E) solids and liquids

ANSWERS
1) A 10x24=240 240≈248
2) B 8.5/120.0 has 2 sig figs
3) E x (1kg/1000g) x (100^3 cm^3/ 1m^3) = x 10^7
4) A n = 10^{-9} p= 10^{-12} nanos are 1000x **bigger** than picos
5) B
6) C 10) D
7) B 11) C estimate
8) D 12) B
9) A 13) A

AP Chemistry Nomenclature Practice Quiz

NOMENCLATURE QUESTIONS **WILL <u>NOT</u>** BE MULTIPLE CHOICE ON TEST

1) The correct name for SrO is __________.
A) strontium oxide B) strontium hydroxide
C) strontium peroxide D) strontium monoxide
E) strontium dioxide

2) The correct name for SO is __________.
A) sulfur oxide B) sulfur monoxide
C) sulfoxide D) sulfate
E) sulfite

3) The correct name for H_2SO_3 is __________.
A) sulfuric acid
B) sulfurous acid
C) hydrosulfuric acid
D) hydrosulfic acid
E) sulfur hydroxide

7) The correct formula of iron(III) bromide is __________.
A) $FeBr_2$ B) $FeBr_3$ C) FeBr
D) Fe_3Br_3 E) Fe_3Br

8) The formula of ammonium carbonate is __________.
A) $(NH_4)_2CO_3$ B) NH_4CO_2
C) $(NH_3)_2CO_4$ D) $(NH_3)_2CO_3$
E) $N_2(CO_3)_3$

10) Chromium and chlorine form an ionic compound whose formula is $CrCl_3$. The name of this compound is _________.
A) chromium chlorine
B) chromium(III) chloride
C) monochromium trichloride
D) chromium(III) trichloride
E) chromic trichloride

12) The formula for aluminum hydroxide is __________.
A) AlOH B) Al_3OH
C) $Al_2(OH)_3$ D) $Al(OH)_3$ E) Al_2O_3

13) The name of the ionic compound $(NH_4)_3PO_4$ is __________.
A) ammonium phosphate
B) nitrogen hydrogen phosphate
C) tetrammonium phosphate
D) ammonia phosphide
E) triammonium phosphate

15) What is the molecular formula for propane __________?
A) C_2H_8 B) C_3H_6 C) C_3H_8
D) C_4H_8 E) C_4H_{10}

16) Which formula/name pair is incorrect?
A) $Mn(NO_2)_2$ manganese(II) nitrite
B) $Mg(NO_3)_2$ magnesium nitrate
C) $Mn(NO_3)_2$ manganese(II) nitrate
D) Mg_3N_2 magnesium nitrite

17) Which species below is the nitride ion?
A) Na^+ B) NO_3^-
C) NO_2^- D) NH_4^+ E) N^{3-}

18) Which species below is the sulfite ion?
A) SO_2^- B) SO_3^-
C) S^{2-} D) H_2SO_4 E) H_2S

19) Which species below is the nitrate ion?
A) NO_2^- B) NH_4^+
C) NO_3^- D) N_3^- E) N^{3-}

20) Which formula/name pair is incorrect?
A) $FeSO_4$ iron(II) sulfate
B) $Fe_2(SO_3)_3$ iron(III) sulfite
C) FeS iron(II) sulfide
D) $FeSO_3$ iron(II) sulfite
E) $Fe_2(SO_4)_3$ iron(III) sulfide

21) Which one of the following is the formula of hydrochloric acid?
A) $HClO_3$ B) $HClO_4$
C) HClO D) HCl
E) $HClO_2$

22) The suffix -ide is used __________.
A) for monatomic anion names
B) for polyatomic cation names
C) for the name of the first element in a molecular compound
D) to indicate binary acids
E) for monoatomic cations

23) The formula for the compound formed between aluminum ions and phosphate ions is __________.
A) $Al_3(PO_4)_3$ B) $AlPO_4$
C) $Al(PO_4)_3$ D) $Al_2(PO_4)_3$ E) AlP

24) What is the name of an alcohol derived from hexane__________?

25) Which metal does not require to have its charge specified in the names of ionic compounds it forms?
A) Mn B) Fe
C) Cu D) Ca E) Pb

Nomenclature Answers

1) A		13) A
2) B		14) C
3) B		15) C
4) C		16) D
5) A		17) E
6) D		18) A
7) B		19) C
8) A		20) E
9) A		21) D
10) B		22) A
11) B	same groups, so same ratio	23) B
12) D		24) hexanol
		25) D

AP Chemistry Chapter Two Practice Quiz

- Know the response of alpha, beta , and gamma radiation in an electric field. What do these responses tell us about the mass and charge of the particles?
- Know the mass, charge, and location of the three subatomic particles (protons, neutrons, and electrons)
- Periodic table (locate and define --group, period, atomic number, atomic mass, metals, non-metals, metalloids, transition metals, alkali metals, alkaline earth metals, chalcogens, halogens, noble gases)
- Development of modern atomic theory --Know how the experiments of Rutherford, Millikan, & Thompson provide evidence of subatomic particles.

1) A certain mass of carbon reacts with 13.6 g of oxygen to form carbon monoxide. __________ grams of oxygen would react with that same mass of carbon to form carbon dioxide, according to the law of multiple proportions?
A) 25.6 B) 6.8 C) 13.6
D) 136 E) 27.2

2) The nucleus of an atom contains __________.
A) electrons
B) protons, neutrons, and electrons
C) protons and neutrons
D) protons and electrons
E) protons

3) Which pair of atoms constitutes a pair of isotopes of the same element?
A) $^{14}_{6}X$ $^{14}_{7}X$ B) $^{14}_{6}X$ $^{12}_{6}X$
C) $^{17}_{9}X$ $^{17}_{8}X$ D) $^{19}_{10}X$ $^{19}_{9}X$
E) $^{20}_{10}X$ $^{21}_{11}X$

4) The element __________ is the most similar to strontium in chemical and physical properties.
A) Li B) At
C) Rb D) Ba E) Cs

5) Horizontal rows of the periodic table are known as ______.
A) periods B) groups
C) metalloids D) metals E) nonmetals

6) Vertical columns of the periodic table are known as _____.
A) metals B) periods
C) nonmetals D) groups E) metalloids

7) Elements in Group 1A are known as the __________.
A) chalcogens B) alkaline earth metals
C) alkali metals D) halogens E) noble gases

8) Potassium is a __________ and chlorine is a __________.
A) metal, nonmetal B) metal, metal
C) metal, metalloid D) metalloid, nonmetal
E) nonmetal, metal

9) _____ are found uncombined, as monatomic species in nature.
A) Noble gases B) Chalcogens
C) Alkali metals D) Alkaline earth metals
E) Halogens

10) When a metal and a nonmetal react, the __________ tends to lose electrons and the __________ tends to gain electrons.
A) metal, metal B) nonmetal, nonmetal
C) metal, nonmetal D) nonmetal, metal
E) None of the above, these elements share electrons.

11) The empirical formula of a compound with molecules containing 12 carbon atoms, 14 hydrogen atoms, and 6 oxygen atoms is __________.
A) $C_{12}H_{14}O_6$ B) CHO
C) CH_2O D) $C_6H_7O_3$ E) C_2H_4O

12) What is the formula of the compound formed between strontium ions and nitrogen ions?
A) SrN B) Sr_3N_2
C) Sr_2N_3 D) SrN_2 E) SrN_3

13) The formula of a salt is XCl_2. The X-ion in this salt has 28 electrons. The metal X is __________.
A) Ni B) Zn
C) Fe D) V E) Pd

14) The charge on the manganese in the salt MnF_3 is __________.

A) +1 B) -1 C) +2
D) -2 E) +3

15) Aluminum reacts with a certain nonmetallic element to form a compound with the general formula AlX. Element X is a diatomic gas at room temperature. Element X must be ____.

A) oxygen B) fluorine
C) chlorine D) nitrogen E) sulfur

16) Potassium forms an ion with a charge of __________.

A) +2 B) -1
C) +1 D) -2 E) 0

17) Calcium forms an ion with a charge of __________.

A) -1 B) -2
C) +1 D) +2 E) 0

18) Iodine forms an ion with a charge of __________.

A) -7 B) +1
C) -2 D) +2 E) -1

19) Predict the empirical formula of the ionic compound that forms from sodium and fluorine.

A) NaF B) Na_2F
C) NaF_2 D) Na_2F_3 E) Na_3F_2

20) Predict the empirical formula of the ionic compound that forms from magnesium and fluorine.

A) Mg_2F_3 B) MgF
C) Mg_2F D) Mg_3F_2
E) MgF_2

21) The ions Ca^{2+} and PO_4^{3-} form a salt with the formula ___.

A) $CaPO_4$ B) $Ca_2(PO_4)_3$
C) Ca_2PO_4 D) $Ca(PO_4)_2$ E) $Ca_3(PO_4)_2$

22) Magnesium and sulfur form an ionic compound with the formula __________.

A) MgS B) Mg_2S
C) MgS_2 D) Mg_2S_2
E) Mg_2S_3

23) Consider some postulates of Dalton's atomic theory:

(i) Each element is composed of extremely small particles called atoms.
(ii) Atoms are indivisible.
(iii) Atoms of a given element are identical.
(iv) Atoms of different elements are different and have different properties.

Which of the postulates is(are) no longer valid?

A) (i) and (ii) B) (ii) only
C) (ii) and (iii) D) (iii) only E) (iii) and (iv)

24) Which pair of substances could be used to illustrate the law of multiple proportions?

A) SO_2, H_2SO_4 B) CO, CO_2
C) H_2O, O_2 D) CH_4, $C_6H_{12}O_6$
E) NaCl, KCl

25) The charge on an electron was determined in the ________.

A) cathode ray tube, by J. J. Thompson
B) Rutherford gold foil experiment
C) Millikan oil drop experiment
D) Dalton atomic theory
E) atomic theory of matter

26) All atoms of a given element have the same __________.

A) mass
B) number of protons
C) number of neutrons
D) number of electrons and neutrons
E) density

27) Which atom has the smallest number of neutrons?

A) carbon-14 B) nitrogen-14
C) oxygen-16 D) fluorine-19
E) neon-20

28) There are __________ electrons, __________ protons, and __________ neutrons in an atom of ${}^{132}_{54}Xe$.

A) 132, 132, 54
B) 54, 54, 132
C) 78, 78, 54
D) 54, 54, 78
E) 78, 78, 132

29) Which isotope has 45 neutrons?

A) ${}^{80}_{36}\mathbf{Kr}$ B) ${}^{80}_{35}\mathbf{Br}$ C) ${}^{78}_{34}\mathbf{Se}$ D) ${}^{34}_{17}\mathbf{Cl}$ E) ${}^{103}_{45}\mathbf{Rh}$

30) Which isotope has 36 electrons in an atom?

A) ${}^{80}_{36}\mathbf{Kr}$ B) ${}^{80}_{35}\mathbf{Br}$ C) ${}^{78}_{34}\mathbf{Se}$ D) ${}^{34}_{17}\mathbf{Cl}$ E) ${}^{36}_{80}\mathbf{Hg}$

31) The atomic mass unit is presently based on assigning an exact integral mass (in amu) to an isotope of __________.

A) hydrogen B) oxygen
C) sodium D) carbon E) helium

32) Vanadium has two naturally occurring isotopes, ${}^{50}V$ with an atomic mass of 49.9472 amu and ${}^{51}V$ with an atomic mass of 50.9440. The atomic weight of vanadium is 50.9415. The percent abundances of the vanadium isotopes are __________% ${}^{50}V$ and __________% ${}^{51}V$.

A) 0.2500, 99.750 B) 99.750, 0.2500
C) 49.00, 51.00 D) 1.000, 99.000
E) 99.000, 1.000

33) An unknown element is found to have three naturally occurring isotopes with atomic masses of 35.9675 (0.337%), 37.9627 (0.063%), and 39.9624 (99.600%). Which of the following is the unknown element?
A) Ar
B) K
C) Cl
D) Ca
E) None of the above could be the unknown element.

34) Of the following, only __________ is not a metalloid.
A) B
B) Al
C) Si
D) Ge
E) As

35) The elements in groups 1A, 6A, and 7A are called, __________, respectively.
A) alkaline earth metals, halogens, and chalcogens
B) alkali metals, chalcogens, and halogens
C) alkali metals, halogens, and noble gases
D) alkaline earth metals, transition metals, and halogens
E) halogens, transition metals, and alkali metals

36) Of the choices below, which one is not an ionic compound?
A) PCl_5
B) $MoCl_6$
C) RbCl
D) $PbCl_2$
E) NaCl

37) Of the following, __________ contains the greatest number of electrons.
A) P^{3+}
B) P
C) P^{2-}
D) P^{3-}
E) P^{2+}

38) There are __________ protons, __________ neutrons, and __________ electrons in $^{131}I^-$.
A) 131, 53, and 54
B) 131, 53, and 52
C) 53, 78, and 54
D) 53, 131, and 52
E) 78, 53, and 72

39) Which metal does not form cations of differing charges?
A) Na
B) Cu
C) Co
D) Fe
E) Sn

Chapter 2 Practice Quiz Answers

1) E
2) C
3) B
4) D
5) A
6) D
7) C
8) A
9) A
10)C
11)D
12)B
13)B
14)E
15)D
16)C
17)D
18)E
19)A
20)E
21)E
22)A
23)C
24)B
25)C
26)B
27)B
28)D
29)B
30)A
31)D
32)A
33)A
34)B
35)B
36)A
37)D
38)C
39)A