CHAPTER 2
ATOMS, MOLECULES, AND IONS

This chapter is a review of first year chemistry concepts and terminology used to describe atoms and molecules, atomic theory, the periodic table, and the naming of compounds.

> **Take Note:** *You need to know and use correctly the basic vocabulary presented in this chapter. A detailed, accurate knowledge of naming rules for compounds is essential for communicating chemical information, both in writing equations and solving problems. Clear and precise communication is essential for successful performance on the AP exam.*

You can organize Chapter 2 material into categories to help you master the information. The main categories are:

- *Key terms pertaining to atoms and compounds*. These terms include: allotrope, atomic number, diatomic elements, molecule, compound, isotope, ion, mass number, law of definite composition, law of multiple proportions, and law of conservation of matter (Table 1).
- *Key terms pertaining to the periodic table*. These terms include: alkali metals, alkaline earth metals, halogens, metals, metalloids, nonmetals, noble gases, family, and period (Table 2).
- *Key scientists and their contributions to atomic theory*. These include: John Dalton, J. J. Thomson, Marie Curie, R. A. Millikan, E. Rutherford, and J. Chadwick (Table 3).
- Rules for naming inorganic compounds.
- Names of common polyatomic ions (Table 4).
- Rules for naming common organic compounds and some examples (Tables 5–9).

Table 1. Key terms pertaining to atoms and compounds

Term	***Definition*** **and some pertinent details**	**Examples**
Allotropes	*Different forms of the same element in the same physical state*.	O_2 and O_3 C diamond and C graphite
Atomic number (*Z*)	The *number of protons (Z)* in the nucleus of an element. The periodic table is organized by increasing atomic number.	Hydrogen $Z = 1$ Helium $Z = 2$
Diatomic elements	Diatomic elements exist as *two-atom molecules.* There are seven elements that exist as diatomic molecules at room conditions.	H_2, O_2, N_2, F_2, Cl_2, Br_2, I_2

Molecule	*Two or more atoms* form a molecule.	H_2O, O_2
Compound	*Two or more <u>different</u> atoms* form a compound. A given compound always has a definite composition.	H_2O, $NaCl$, $C_6H_{12}O_6$ O_2 is a molecule but *not* a compound because it only has one type of atom.
Isotope	Isotopes are atoms of the same element that have *different numbers of neutrons*. Because of the different numbers of neutrons, isotopes of the same element have different mass numbers and some different properties. For example, the isotope C-12 has a mass number of 12 and is not radioactive. The isotope C-14 has a mass number of 14 and is radioactive.	C-12 and C-14
Ion	An ion is a *charged species* in which the number of protons does not equal the number of electrons. A **cation** is a positive ion; an **anion** is a negative ion. **Monatomic ions** contain one atom and **polyatomic ions** contain more than one atom.	Monatomic cation: Na^+ Monatomic anion: Cl^- Polyatomic cation: NH_4^- Polyatomic anion: OH^-
Mass number (*A*)	The mass number is equal to the *number of protons plus neutrons* in the nucleus. The mass number minus the number of protons gives the number of neutrons.	U-235 refers to the uranium isotope with 92 protons and a mass number of 235.
Law of Definite Composition	*A given compound always has the same composition.*	Water is always made up of two hydrogen atoms and one oxygen atom. Table salt is always made up of one sodium atom and one chloride atom.
Law of Multiple Proportions	Sometimes *two elements can combine in more than one way* to form two or more different compounds. Subscripts are simple whole number multiples of each other.	H_2O and H_2O_2 In H_2O, 2 g H combine with 16 g O (1 to 8 ratio). In H_2O_2, 2 g H combine with 32 g O (1 to 16 ratio).
Law of Conservation of Matter	*Matter can neither be created nor destroyed* in a nonnuclear chemical reaction.	Mass of reactant molecules = mass of product molecules. $2H_2 + O_2 \rightarrow 2H_2O$ $2(2\ g) + 32\ g \rightarrow 2(18\ g)$ $36\ g \rightarrow 36\ g$

Table 2. Key terms pertaining to the periodic table

Term	Definition and some pertinent details
Alkali metals	Group 1A elements: Li, Na, K, Rb, Cs, and Fr. They are the most active metals on the periodic table. They lose one electron and form 1+ ions.
Alkaline Earth Metals	Group 2A elements: Be, Mg, Ca, Sr, Ba, and Ra. They form 2+ ions.
Halogens	Group 7A elements: F, Cl, Br, I, and At. In binary compounds they form 1– ions.
Metals	The metals are on the left side and center of the periodic table. Metals are shiny, malleable, ductile, and are good conductors of heat and electricity.
Metalloids	These elements (B, Si, Ge, As, Sb, Te, Po, At) have properties intermediate between a metal and a nonmetal.
Nonmetals	Nonmetals are on the right side of the periodic table and are poor conductors of heat and electricity.
Noble gases	Group 8A elements: He, Ne, Ar, Kr, Xe, and Rn. He, Ne, and Ar are not chemically active under any conditions.
Family	A vertical column in the periodic table; also known as a **group**. Members of a group have the same number of valence electrons (electrons in the outermost energy shell) and somewhat similar chemical properties.
Period	A horizontal row in the periodic table; also known as a **series**.

Table 3. Key scientists and their contributions to atomic theory

Year	Scientist	Scientific contribution
1808	**John Dalton**	Dalton's atomic theory: • All matter is composed of atoms. • Atoms of one element are all identical in mass and chemical behavior. • Atoms are neither created nor destroyed in chemical reactions. • Atoms combine in small whole number ratios to form compounds. • A given compound has a constant composition (number and type of atoms are the same).
1897	**J. J. Thomson**	Using a cathode ray tube, Thomson determined that cathode rays were composed of negatively charged particles (later called electrons). Thomson calculated the charge-to-mass ratio for electrons. Thomson's model of the atom (the 'plum pudding' model) was later disproved by Rutherford.

1900	**Marie Curie**	Curie worked on the nature of radioactivity and discovered polonium and radium.
1909	**R. A. Millikan**	Millikan performed the oil drop experiment, which enabled him to calculate the charge of an electron.
1911	**E. Rutherford**	Rutherford performed the gold foil experiment in which gold foil was bombarded with alpha particles (helium nuclei). Most alpha particles went through the foil undeflected, leading to the conclusion that the atom was mostly empty space with a very tiny, dense, and positive center (the nuclear model of the atom). Rutherford also studied the nature of radioactivity and described the three types of radiation: alpha, beta, and gamma radiation.
1932	**J. Chadwick**	Chadwick discovered the neutron.

Rules for naming inorganic compounds

1. **Naming a binary compound with a metal and a nonmetal**. By convention, the symbol for the metal is written first. The total charge of a compound is zero. A **binary compound**, a compound *composed of two elements*, ends with the suffix -*ide*. An example is NaCl, sodium chloride.

2. **Naming binary metal/nonmetal compounds in which the metal has more than one oxidation state**. Common metal ions with multiple valences are:
   a. Cu^+, Cu^{2+} named copper(I) and copper(II) or cuprous and cupric ion
   b. Fe^{2+}, Fe^{3+} named iron(II) and iron(III) or ferrous and ferric ion
   c. Sn^{2+}, Sn^{4+} named tin(II) and tin(IV) or stannous and stannic ion
   d. Hg_2^{2+}, Hg^{2+} named mercury(I) and mercury(II) or mercurous and mercuric ion. Note that the mercury(I) ion occurs as the dimer, Hg_2^{2+}.

Some helpful hints to help you with your memorization of oxidation numbers:

a. The elemental state is zero (Na^0, K^0).
b. Family 1A metal ions have a +1 oxidation state (Na^+, K^+).
c. Family 2A metal ions have a +2 oxidation state (Mg^{2+}, Ba^{2+}).
d. Family 7A (halogen) ions have a –1 oxidation state in binary compounds (KCl, HF).
e. Common cations that have only one oxidation state are: Al^{3+}, Zn^{2+}, and Ag^+.
f. Hydrogen usually has a +1 oxidation state (H^+) unless it is a hydride in which hydrogen has a –1 oxidation state (H^-). Hydrides form between hydrogen and a very active metal; an example is sodium hydride, NaH.
g. Oxygen usually has an oxidation state of –2 in oxides (O^{2-}) unless it is a peroxide in which oxygen has a –1 oxidation state (O_2^{2-}). For example, Na_2O and H_2O_2 are called sodium oxide and hydrogen peroxide, respectively.
h. See Figure 2.11 for some examples of common monatomic ions.

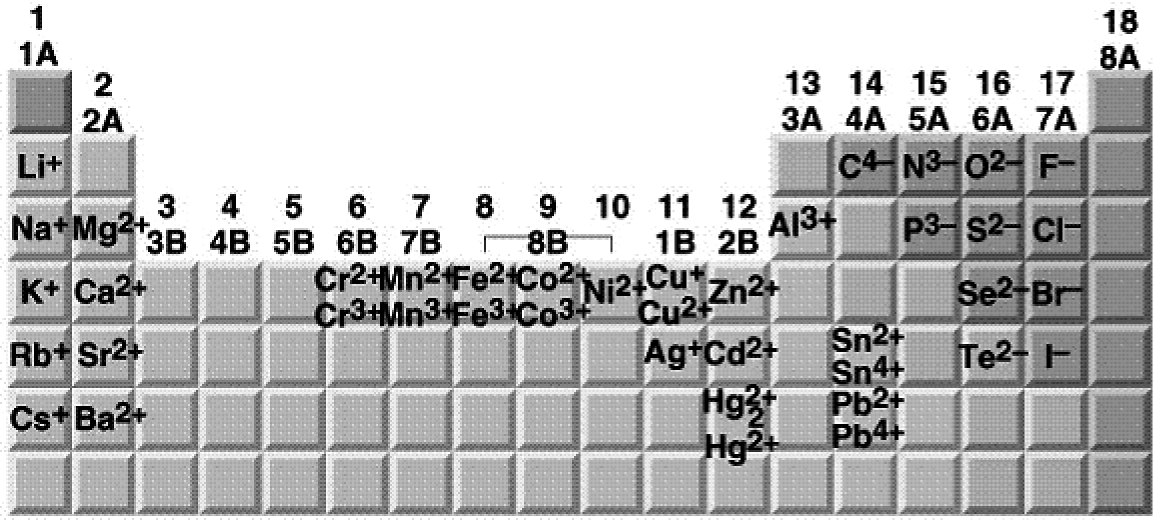

Figure 2.11

3. **Naming binary compounds with two nonmetals (molecular compounds).** Prefixes of mono-, di-, tri-, etc., are necessary when one pair of elements can form several different molecular compounds. Elements such as C, N, and S have multiple valences and the name must identify which oxidation number is present. Common examples are:
    a. CO and CO_2 are named carbon monoxide and carbon dioxide, respectively. Carbon oxide is an incorrect name because it does not distinguish between the two compounds.
    b. SO_2 and SO_3 are named sulfur dioxide and sulfur trioxide, respectively.
    c. CF_4 is named carbon tetrafluoride. C usually has an oxidation state of +4 or −4. Refer to Tables 5 through 9 for more organic molecule naming guidelines.
    d. NO and N_2O_4 are named nitrogen monoxide and dinitrogen tetraoxide, respectively.

4. **Naming compounds that contain a metal and a polyatomic ion**. When compounds contain polyatomic ions, simply name the metal and name the polyatomic ion. You do not make any changes to the name of the metal ion or the polyatomic ion. For example, $NaNO_3$ is simply sodium nitrate.

Table 4. Most frequently used polyatomic ions

Acetate	CH_3COO^-
Ammonium	NH_4^+
Carbonate	CO_3^{2-}
Chlorate*	ClO_3^-
Chromate	CrO_4^{2-}
Dichromate	$Cr_2O_7^{2-}$
Hydrogen sulfate	HSO_4^-
Nitrate*	NO_3^-
Permanganate	MnO_4^-
Phosphate*	PO_4^{3-}
Sulfate*	SO_4^{2-}

* Refers to the most common form of the polyatomic ion. The number of oxygen atoms can vary with these polyatomic ions but the charge of the polyatomic ion remains the same. Check the naming rules below for the less common forms of these polyatomic ions.

Parentheses are used when the polyatomic ion is used more than once in a compound. For example, barium chlorate is written as $Ba(ClO_3)_2$. If the polyatomic ion is only used once in a compound, the parentheses are not used. For example, sodium chlorate is written as $NaClO_3$.
Some polyatomic ions exist in more than one form:

a. The name for the most common form of a polyatomic ion ends in *-ate*. Examples are: ClO_3^- chlorate, NO_3^- nitrate, PO_4^{3-} phosphate, and SO_4^{2-} sulfate.
b. If one oxygen is added to the most common form of the polyatomic ion a prefix of *per-* and suffix of *-ate* are used. For example, ClO_4^- is called perchlorate. Not all polyatomic ions can add an extra oxygen.
c. If one oxygen is removed from the most common form the name ends in *-ite*. For example, ClO_2^- is called chlorite.
d. If more than one oxygen is removed from the most common form a prefix of *hypo-* and a suffix of *-ite* are used. For example, ClO^- is called hypochlorite.

Take Note: *If you learn the most common form of the polyatomic ion and the rules above in a to d, you can deduce the names of most polyatomic ions that will appear on the AP exam.*

5. **Naming acids.** (Most acids have H listed as the first element in the compound.)

    a. **Binary acids** (*made of two elements*). The name is composed of prefix *hydro-*, stem and suffix of *-ic*. Examples are HCl and HI named *hydro*chlor*ic* acid and h*ydro*iod*ic* acid, respectively.

    b. **Oxoacids** (made of hydrogen and a polyatomic ion that contains oxygen). If the polyatomic ion ends in *-ate*, replace the *-ate* ending with *-ic acid*. Examples are:

        ClO_3^- chlorate becomes chloric acid ($HClO_3$)

        ClO_4^- perchlorate becomes perchloric acid ($HClO_4$)

        ClO_2^- chlorite becomes chlorous acid ($HClO_2$)

        ClO^- hypochlorite becomes hypochlorous acid ($HClO$)

    Note that prefix and suffix rules mirror the polyatomic ion rules given in 4.

6. **Naming bases**. For simple bases containing the –OH group the name is composed of the metal name plus the word *hydroxide.* Examples are NaOH and $Al(OH)_3$, named sodium hydroxide and aluminum hydroxide, respectively.

7. **Naming hydrates**. The hydrate name is made up of several components listed in the following order: metal name, anion name, prefix to indicate the number of water molecules, followed by the word *hydrate.* An example is $CuSO_4 \cdot 5H_2O$ named copper(II) sulfate pentahydrate.

8. **Common and systematic names of some compounds** are given in Table 2.7.

TABLE 2.7 Common and Systematic Names of Some Compounds

Formula	Common Name	Systematic Name
H_2O	Water	Dihydrogen monoxide
NH_3	Ammonia	Trihydrogen nitride
CO_2	Dry ice	Solid carbon dioxide
NaCl	Table salt	Sodium chloride
N_2O	Laughing gas	Dinitrogen monoxide
$CaCO_3$	Marble, chalk, limestone	Calcium carbonate
CaO	Quicklime	Calcium oxide
$Ca(OH)_2$	Slaked lime	Calcium hydroxide
$NaHCO_3$	Baking soda	Sodium hydrogen carbonate
$Na_2CO_3 \cdot 10H_2O$	Washing soda	Sodium carbonate decahydrate
$MgSO_4 \cdot 7H_2O$	Epsom salt	Magnesium sulfate heptahydrate
$Mg(OH)_2$	Milk of magnesia	Magnesium hydroxide
$CaSO_4 \cdot 2H_2O$	Gypsum	Calcium sulfate dihydrate

Rules for naming organic compounds and some common examples.

Take Note: *You will be required to know the names of simple organic compounds and recognize common functional groups (Tables 5 to 9) on the AP exam.*

Table 5. Names of the first ten alkanes (saturated hydrocarbons, single bonds only, name ends in -*ane*, homologous formula: C_nH_{2n+2})

Formula	Name
CH_4	Methane
C_2H_6	Ethane
C_3H_8	Propane
C_4H_{10}	Butane
C_5H_{12}	Pentane
C_6H_{14}	Hexane
C_7H_{16}	Heptane
C_8H_{18}	Octane
C_9H_{20}	Nonane
$C_{10}H_{22}$	Decane

Table 6. Names of alkenes (unsaturated hydrocarbons, contain a double bond, name ends in -*ene*, homologous formula: C_nH_{2n})

C_2H_4	Ethene (common name is ethylene)
C_3H_6	Propene
C_4H_8	Butene

Table 7. Names of alkynes (unsaturated hydrocarbons, contain a triple bond, name ends in -*yne*, homologous formula: C_nH_{2n-2})

C_2H_2	Ethyne (common name is acetylene)
C_3H_4	Propyne
C_4H_6	Butyne

Table 8. Common functional groups

Functional group	Example	Name
R – OH Alcohol	C_2H_5OH	Ethyl alcohol (ethanol)
R – COOH Carboxylic Acid	CH_3COOH	Acetic acid (ethanoic acid)
R – COOR Ester	CH_3COOCH_3	Methyl acetate
R – CHO Aldehyde	CH_3CH_2CHO	Propanal
R – O – R Ether	CH_3OCH_3	Dimethyl ether
O || R — C — R Ketone	O || CH_3 — C — CH_3	Dimethyl ketone (acetone)

Table 9. Simple organic compounds that are often mentioned

Molecular fomula	Name
C_6H_6	Benzene
$C_6H_{12}O_6$	Glucose
C_2H_2	Acetylene
C_2H_4	Ethylene

SAMPLE MULTIPLE CHOICE QUESTIONS

Questions 1 – 4 refer to the following elements.

A. Br
B. C
C. P
D. Zn
E. Rn

1. The element is a member of the noble gas family.
2. The element exists as a diatomic molecule at room conditions.
3. The element forms Na_2XO_3 compound.
4. The element has a monatomic ion with a charge of 2+.

5. Which of the following pairs are isotopes?

   A. O_2 and O_3
   B. Cu^+ and Cu 2^+
   C. U-235 and U-238
   D. Hydrogen and helium
   E. C_2H_2 and CH

6. Which of the following compounds is incorrectly named?

   A. $NaHSO_4$ sodium hydrogen sulfate
   B. $K_2Cr_2O_7$ potassium dichromate
   C. Fe_2O_3 iron(II) oxide
   D. CO carbon monoxide
   E. $Ba(NO_2)_2$ barium nitrite

7. Which family represents the most active metals?

   A. Alkali metals
   B. Alkaline earth metals
   C. Transitional metals
   D. Family 7A, the halogens
   E. Family 8A, the noble gases

8. The scientist who developed the nuclear model of the atom is

   A. John Dalton
   B. J. J. Thomson
   C. Ernest Rutherford
   D. R. A. Millikan
   E. Albert Einstein

9. Which polyatomic ion is incorrectly named?

   A. PO_4^{3-} phosphate
   B. SO_4^{2-} sulfite
   C. CrO_4^{2-} chromate
   D. CO_3^{2-} carbonate
   E. NH_4^+ ammonium

10. The empirical formula for glucose, $C_6H_{12}O_6$, is

   A. CH_2
   B. CH_2O
   C. $C_3H_6O_3$
   D. CHO
   E. None of these

<u>Comprehension Questions</u>

1) Differentiate among atoms, ions, molecules, and compounds using the following substances: Ne, F_2, PCl_5, $KClO_3$, and SO_4^{2-}. Be sure to comment on the number and type of atom(s) involved, electrical charge, types of elements (metal or nonmetal) involved, and characteristics of the formula that allow for identification and differentiation between species.

2) Define the term *empirical.* How does its meaning relate to the term *empirical formula?* Discuss the differences and similarities between a molecular formula and empirical formula.

3) Provide names for the following compounds and ions: H_2SO_4, SO_2, NO_2^-, Li_3PO_4, KBr, CO_2, and I^-.

4) Provide formulae for the following compounds: hypochlorite, strontium ion, phosphorus trichloride, hydrobromic acid, and potassium hydrogen carbonate.

ANSWERS TO SAMPLE MULTIPLE CHOICE QUESTIONS

1. E

   The noble gas family is Family 8A (last column) on the periodic table. Other common groups to know: Family 1A alkali metals, Family 2A alkaline earth metals, Family 7A halogens.

2. A

   Bromine, Br_2, is a diatomic element. The other elements that exist as diatomic molecules at room conditions are: H_2, N_2, O_2, F_2, Cl_2, I_2.

3. B

   Na_2CO_3. This is the formula for sodium carbonate. It is important to know all the polyatomic ions listed in Table 4.

4. D

   Zinc forms 2+ ion. This is the only oxidation state for zinc.

5. C

U-235 and U-238 are isotopes. They have the same number of protons (92) but have a different number of neutrons (143 and 146, respectively).

6. C

The correct name for Fe_2OH_3 is iron(III) oxide. The iron has a +3 oxidation state in this compound. The Roman numeral is used to characterize the oxidation state of the metal ion when the metal has numerous oxidation states. Some common metal ions where using the Roman numeral system is applicable are : $Cu^{+, 2+}$; Hg_2^{2+} / Hg^{2+}; $Fe^{2+, 3+}$; $Sn^{2+, 4+}$.

7. A

The alkali metals are the most active metals. When forming compounds they always have an oxidation state of +1.

8. C

Rutherford's gold foil experiment led to the model of the nuclear atom. In this experiment alpha particles (helium nuclei) were beamed at gold foil. Most of the alpha particles passed through the foil undeflected and very few bounced back. This led to the conclusion that the atom was mostly empty space with a tiny, dense, positive center.

9. B

SO_4^{2-} is the sulfate polyatomic ion. Sulfite is SO_3^{2-}.

10. B

The empirical formula is the most reduced form of the molecular formula. $C_6H_{12}O_6$ is the molecular formula and the empirical formula is CH_2O.

<u>Answers to Comprehension Questions</u>

1) An atom can be considered a singular particle of matter, and is the smallest amount of an element that retains the properties of that element. There are about 118 different kinds of atoms known to date, and Ne is an example of a substance that exists as isolated atoms. F_2 is a molecular substance, that is, the particles of this substance consist of two atoms bonded together as opposed to the isolated atoms found in the case of Ne.

Ions are charged particles. They can consist of a single atom or groups of differing or similar atoms. SO_4^{2-} is an example of a polyatomic ion as it is a charged particle consisting of several atoms of different elements bonded together. A compound is a pure substance composed of atoms of different elements. Compounds differ from mixtures in that the elements present in a compound always appear in the same ratios, whereas the elements contained in a mixture have no fixed ratios. PCl_5 and $KClO_3$ are both examples of compounds with fixed ratios of elements. PCl_5 is a molecular compound and $KClO_3$ is an ionic compound. The distinction between these two has to do with the types of bonds holding the individual atoms together to make up the compound.

2) Empirical can be defined as *determined from actual observations or experimental results*. A process known as elemental analysis can be used to determine the mass ratio of elements in a compound. From these mass ratios and atomic masses of the elements a molar ratio of elements can be determined for a compound. This molar ratio is known as an empirical formula as it is derived from experimental results. This formula may or may not represent the molecular formula for a compound depending on how many atoms actually constitute a single molecule or formula unit of the compound. A separate determination of the molecular mass is needed to derive the molecular formula.

3) H_2SO_4, **sulfuric acid**; SO_2, **sulfur dioxide**; NO_2^-, **nitrite**; Li_3PO_4, **lithium phosphate**; KBr, **potassium bromide**; CO_2, **carbon dioxide**; and I^-, **iodide**.

4) Hypochlorite, **ClO^-**; strontium ion, **Sr^{2+}**; phosphorus trichloride, **PCl_3**; hydrobromic acid, **HBr**; and potassium hydrogen carbonate, **$KHCO_3$**.