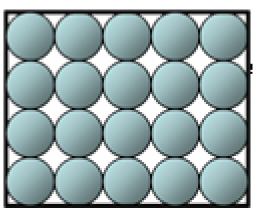
Chapter 1: Atoms, Molecules and Ions

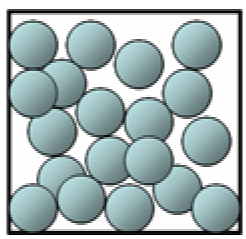
Section 1.1: Introduction

In this course, we will be studying **matter**, “the stuff things are made of”. There are many ways to classify matter. For instance, matter can be classified according to the **phase**, that is, the physical state a material is in. Depending on the pressure and the temperature, matter can exist in one of three phases (solid, liquid, or gas). The chemical structure of a material determines the range of temperatures and pressures under which this material is a solid, a liquid or a gas.

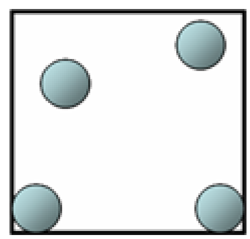
Consider water for example. The principal differences between water in the solid, liquid and gas states are simply: 1) the average distance between the water molecules; small in the solid and the liquid and large in the gas and 2) whether the molecules are organized in an orderly three-dimensional array (solid) or not (liquid and gas).



Solid State
Ordered and dense
Has a definite shape and volume.
Solids are very slightly compressible.



Liquid State
Disordered and usually slightly less dense.
Has a definite volume and takes the shape of the container.
Liquids are slightly compressible.



Gas State
Disordered and much lower density than crystal or liquid.
Does not have definite shape and volume.
Gases are highly compressible.

Another way to classify matter is to consider whether a substance is pure or not. So, matter can be classified as being either a **pure substance** or a **mixture**. <u>A pure substance has unique composition and properties</u>. For example, water is a pure substance (whether from Texas or Idaho, each water molecule always contains 2 atoms of hydrogen for 1 atom of oxygen). Under the same atmospheric pressure and at the same ambient temperature, water always has the same density.

We can go a little further and classify mixtures are either **homogeneous** or **heterogeneous**. In a homogeneous mixture, for example, as a result of mixing a teaspoon of salt in a glass of water, the composition of the various components and their properties are the same throughout. Different aliquots of this salt solution would have the same density. In contrast, dropping gold coins or a teaspoon of oil in a glass of water will result in the formation of a heterogeneous mixture. Different aliquots will contain different amounts of oil or of gold depending on whether these aliquots are taken from the top or the bottom of the mixture. A homogeneous system exhibits a single phase, while a heterogeneous one exhibits multiple phases (different solids, liquids or mixtures of these).

In the rest of this chapter, we will focus on **pure substances**. There are only two kinds of pure substances: 1) **elements** and 2) **compounds**.

<u>Elements are the simplest form of matter and cannot be broken down using chemical methods into two or more pure substances</u>.

> For example, iron is a pure substance, You can take a piece of iron and break it down into smaller and smaller pieces, but each of these smaller pieces has the same properties as the starting material (hence, it is always the same substance).

<u>Compounds, on the other hand, can be broken down into two or more pure substances</u>.

> For example, H_2O or water can be broken down into H_2 (hydrogen gas) and O_2 (oxygen gas). Similarly, table salt or NaCl can be broken down into Na (sodium metal) and Cl_2 (chlorine gas). Compounds are therefore defined as being made of at least two different elements. A compound is a pure substance with unique composition and properties. Hence, NO_2 and N_2O are different compounds since they have different compositions.

Compounds made with only two elements (such as H_2O, NO_2, N_2O, NaCl) are called **binary compounds**.

The chart showing all known elements and giving some of their properties is the **Periodic Table of the Elements**.

Section 1.2: Basic Periodic Table

Practice and memorize the first 54 elements (from H: hydrogen to Xe: xenon) of the Periodic Table. Also memorize the following elements (Cs, Ba, W, Os, Pt, Au, Hg, Pb, Bi) as they appear frequently in many chemical problems and engineering applications.

Section 1.3: Concept of Atomic Number

In 1808, John Dalton developed the **Atomic Model of Matter**. As is the case for every model, Dalton's model is based on a number of assumptions or "postulates".

1) An element is composed of particles called **atoms**. All these atoms exhibit the same chemical properties.

2) In a chemical reaction (transformation of pure substances called reactants into other pure substances called products), no atom of any element is destroyed, created or changed into an atom of another element.

3) Compounds are formed when atoms of two or more elements combine in a unique fashion ($CaCl_2$, H_2O, etc…).

An atom is the smallest "amount" of an element that has the properties of the element. Individual atoms are made of subatomic particles (**electrons**, **protons** and **neutrons**).

Electrons:

1) Are characterized by the symbol $\mathbf{e}^-$

2) Reside in the outer regions of a given atom

3) Have a very small mass, $m_{e-} = 9.11 \times 10^{-28}$ grams (g)

4) Have a negative electric charge $q_{e-} = -1.6 \times 10^{-19}$ coulombs (C)

Protons:

1) Are characterized by the symbol **p**

2) Reside in the central part of the atom, known as the **nucleus**

3) Have a larger mass than the electron, $m_p = 1.673 \times 10^{-24}$ grams (g)

4) Have a positive electric charge $q_p = +1.6 \times 10^{-19}$ coulombs (C)

Neutrons:

1) Are characterized by the symbol **n**

2) Reside in the nucleus

3) Have about the same mass as the protons, $m_n = 1.675 \times 10^{-24}$ grams (g)

4) Have a zero electric charge (i.e. they are neutral)

Note:

1) Protons and neutrons are called **nucleons** as they are all located in the nucleus.

2) To make things look simpler, (which sometimes leads to more confusion), <u>we express charges as multiples of the charge of a proton.</u> Hence, we often say that a proton has a charge of +1 and an electron has a charge of -1 and in this case we do not specify the units of charge. We understand that the unit of charge is the charge of 1 proton (that is:1.6×10^{-19} C).
The key concept to remember is that the charge of the electron is equal in magnitude and opposite in sign to that of the proton.

3) The vast majority of the mass of the atom resides in the nucleus.

All atoms of a given element have the same number of protons in their nucleus. <u>It is the number of protons which characterizes the element</u>. Hence, the number of protons in the nucleus is given the name **atomic number** and is denoted by the symbol **Z**.

Section 1.4: Periodic Table (Atomic Numbers)

Practice with this Interactive Periodic Table and note that the periodic table is built in such a way that elements are placed in order of increasing atomic number from left to right and from top to bottom.

Section 1.5: Concept of Isotope

We now define the **mass number** for elements as the sum of the number of protons and neutrons (remember most of the mass of an atom resides in the nucleus). The mass number is given the symbol **A**.

We can therefore represent an element by its **nuclear symbol**: ${}^{A}_{Z}E$

Since Z is the number of protons and A is the number of protons + neutrons, then: A – Z is the number of neutrons. All atoms of an element must have the same Z values but some atoms may have different A values. Atoms that have the same number of protons but different numbers of neutrons are called **isotopes**.

For instance, ^{12}C, ^{13}C and ^{14}C are three well known isotopes of carbon. They are called carbon-12, carbon-13 and carbon-14. The element hydrogen has 3 isotopes, ^{1}H, ^{2}H (called deuterium) and ^{3}H (called tritium).

All isotopes of an element have the same number of electrons and the same number of protons, since atoms are always neutral. All carbon isotopes have 6 electrons and 6 protons (Z = 6). However, carbon-12 has 6 neutrons, carbon-13 has 7 neutrons, and carbon-14 has 8 neutrons.

Isotopes are generally not present in nature in equal quantities. Hence, we say that isotopes have different natural abundances.

Practice the Interactive Problems to fully understand these concepts.

To learn more about isotopes and their stability, go to Chapter 20, where radioactivity is discussed in detail.

Section 1.6: Periodic Table (Isotopes)

Use this Periodic Table to learn about some of the common isotopes of well known elements.

Section 1.7: Metals, Nonmetals and Metalloids

Elements in the Periodic Table can be classified as **metals**, **nonmetals** and **metalloids** or **semimetals**. Metals are typically on the left-hand side of the Periodic Table (exception: H is a nonmetal). Nonmetals are typically on the right-hand side of the Periodic Table and metalloids on either side of a stairway between metals and nonmetals starting between Boron and Aluminum.

Metals are characterized by the following physical properties:

1) Luster, high heat and electrical conductivity

2) Malleability (ability to make films or sheets)

3) Ductility (i.e. they can be pulled into wires)

Nonmetals do not exhibit the above properties.

Metalloids or semi-metals have some properties of metals and some properties of nonmetals. Metalloids include boron (B), silicon (Si), germanium (Ge), arsenic (As), antimony (Sb), tellurium (Te) and polonium (Po).

Knowing which elements are metals and which are nonmetals is an absolute necessity as far as naming ions and compounds is concerned. Different rules apply for the naming of a compound depending on whether the compound includes metallic elements or not.

Section 1.8: Periodic Table (Metals, Nonmetals and Metalloids)

Review which elements are metals, which are nonmetals and which are metalloids (or semi-metals). Know how to locate the famous “stairway”.

Sections 1.9 - 1.10: Concepts of Group and Period

The horizontal rows in the Periodic Table are called **periods.** The vertical columns in the Periodic Table are called **groups**. There are 18 groups and 7 periods in the Periodic Table.

In Section 1.10, practice with the Interactive Periodic Table to know the location of elements in respective periods and groups.

Sections 1.11 - 1.12: Concept of Family

Some of the groups and some sets of groups have specific names and constitute **families**. Here are the families you need to know:

1) Elements in groups 1, 2, 13, 14, 15, 16, 17, 18 are called **main group elements**.

2) Elements in groups 3 through 12 are called **transition elements**.

3) Elements in group 1, except hydrogen, are called **alkali metals**.

4) Elements in group 2 are called **alkaline earth metals**.

5) Elements in group 17 are called **halogens**.

6) Elements in group 18 are called **noble gases**.

In Section 1.12, practice locating elements and their respective families using the Interactive Periodic Table.

Sections 1.13 - 1.14: Properties of Elements in a Family

Elements in the same family tend to have similar properties. Properties, while similar, may be of a different magnitude.

For example, consider C, Si, Ge. We will see in Chapter 8 and Chapter 9 that they exhibit similar bonding with atoms like H or Cl (i.e. form molecules of similar shape).

Consider Br, Cl, I: They react by a similar mechanism with hydrocarbons (molecules obtained from crude oil that contain only C and H).

Consider Li, Na and K. They react by a similar mechanism with water, (see video in Section 1.14).

Understanding why elements in the same family have related properties is discussed in Chapter 7.

Section 1.15: Concept of Molecule (Part I)

Molecules, like all compounds, are neutral (no net charge). Molecules form when two or more atoms of the same or of different nonmetal elements combine with one another. By "combine" we mean that they form **chemical bonds** between them.

There are principally two types of chemical bonds:

1) **Ionic bonds** are chemical bonds between a metal and a nonmetal. For example NaCl, CsF, PbCl are ionic bonds.

2) **Covalent bonds** are chemical bonds between two nonmetals. For example CH, NO, HCl, CO, SO, PCl are covalent bonds.

Note: A significant number of atoms in the Periodic Table exist under normal conditions in the "**elemental form**" as solids. For example, Fe (iron) exists as an element as solid iron. Similarly C (carbon) exists as solid graphite or solid diamond. There are however a number of notable exceptions that you need to be aware of.

Under normal conditions (atmospheric pressure and ambient temperature):

1) He, Ne, Ar, Kr, Xe, Rn or noble elements exist as gases.

2) H, N, O, F, Cl are not stable in the elemental form and exist as gases H_2, N_2, O_2, F_2, Cl_2 in the "**molecular form**". Hence, when we say hydrogen gas, oxygen gas, nitrogen gas, etc… we always refer to H_2, N_2, O_2 etc…not H, N, O.

3) Br and I are not stable in the elemental form and exist as liquid Br_2 and solid I_2.

4) Phosphorus, P, and sulfur, S, are not stable in the elemental form and exist as P_4 and S_8 in the molecular form.

5) Hg (a metal) in the elemental form exists as a liquid.

Section 1.16: Periodic Table (Molecules)

Interact with the Periodic Table to learn that H, N, O, F, Cl, Br, I, P and S exist under normal conditions in the molecular form (H_2, N_2, O_2, F_2, Br_2, I_2, P_4 and S_8).

Section 1.17: Concept of Ion (Part I)

When an atom loses or gains electrons, a charged particle is formed. This charged particle is called an **ion**.

Typically, metal elements tend to lose electrons, forming positively charged ions called **cations**.

For example:

$_{11}Na \rightarrow {}_{11}Na^+ + 1\ e^-$

(Note: group 1 metals always lose 1 e^- when forming ions)

Atom	Cation
11 p	11 p
11 e^-	10 e^-

$_{20}Ca \rightarrow {}_{20}Ca^{+2} + 2\ e^-$

(Note: group 2 metals always lose 2 e^- when forming ions)

Atom	Cation
20 p	20 p
20 e^-	18 e^-

Note: In many modern text books the ion Ca^{+2} is expressed as Ca^{2+}. However, the authors of this DVD text have decided to express ions in general as ion$^{+(number)}$, as it leads to less confusion when hearing the formula of a polyatomic ion.

Typically, nonmetal elements tend to gain electrons, forming negatively charged ions called **anions**.

For example:

$_{17}Cl + 1\ e^- \rightarrow {}_{17}Cl^-$

(Note: halogens always gain 1 e^- when forming ions)

Atom	Anion
17 p	17 p
17 e^-	18 e^-

$_8O + 2\ e^- \rightarrow {}_8O^{-2}$

(Note: in general, group 16 elements gain 2 e^- when forming ions)

Atom	Anion
8 p	8 p
8 e^-	10 e^-

Note: In many modern text books the ion S^{-2} is expressed as S^{2-}. However, the authors of this DVD text have decided to express ions in general as ion$^{-(number)}$, as it leads to less confusion when hearing the formula of a polyatomic ion.

Practice with the Interactive Problems to fully understand these concepts.

Section 1.18: Periodic Table (Ions)

Interact with the Periodic Table, by clicking on the group numbers highlighted in red. Memorize the types of ions that are predicted to form for elements in different groups. Remember that it is the group an element belongs to, which matters as far as forming ions is concerned.

Section 1.19: Concept of Ion (Part II)

We learned that:

1) Group 1 elements (Li, Na, K, etc...) tend to lose 1 electron

2) Group 2 elements (Be, Mg, Ca, Ba, etc…) tend to lose 2 electrons.

3) Group 13 metals (Al, Ga, In) tend to lose 3 electrons.

4) Group 15 nonmetals tend to gain 3 electrons.

5) Group 16 nonmetals tend to gain 2 electrons.

6) Group 17 nonmetals (halogens) tend to gain 1 electron, forming **halides.**

7) Group 18 (noble gases) do not form stable ions (they are mostly inert).

Why is this? Because "**Nobility is Stability**"

What do we mean by Nobility is Stability?

Consider Li. Lithium has 3 electrons and 3 protons. Lithium is in group 1. Hence, it forms a stable ion (Li^+) by losing 1 electron. Hence, Li^+ has 3 protons and 2 electrons. *Li^+ has the same number of electrons as Helium (He)*

Consider Mg. Magnesium has 12 protons and 12 electrons. Magnesium is in group 2. Hence, it forms a stable ion (Mg^{+2}) by losing 2 electrons. Hence, Mg^{+2} has 12 protons and 10 electrons. *Mg^{+2} has the same number of electrons as Neon (Ne)*.

Consider Br. Bromine has 35 electrons and 35 protons. Bromine is in group 17. Hence, it forms a stable ion (Br^-) by gaining 1 electron. Br^- has 35 protons and 36 electrons. *Br^- has the same number of electrons as Krypton (Kr).*

Stable ions are often formed through a gain or a loss of electrons that is such that the resulting ions have the same number of electrons as the closest noble gas element in the Periodic Table. This statement is rigorously true for all metals in groups 1, 2, for the light metals in group 13 (Al, Ga) and for all nonmetals in groups 15, 16 and 17. Note that transition metals (groups 3 through 12) and metals in group 14 do not always obey this rule.

Overview of Sections 1.20 - 1.51: Naming of Compounds

In the remainder of Chapter 1, we review the rules used for the naming of compounds. Again, recall that there are two main types of compounds: ionic and molecular (or covalent). Different rules apply for ionic and for molecular compounds. First, we discuss ionic compounds.

Ionic compounds are formed by combination of cations with anions in such a way that the resulting substance has a formula that is electrically neutral. For example, Na^{+} combines with Cl^{-} to form NaCl (overall electrically neutral). Ca^{+2} combines with 2 F^{-} to form CaF_2 (also electrically neutral). Depending on the cation and anion types, different naming schemes are used. The different groups of cations (left-hand side) and anions (right-hand side) are:

Cations	**Anions**
Main Group Metal Cations	
Transition Metal Cations	Nonmetal Anions
Polyatomic Cation (NH_4^{+})	Polyatomic Anions
Hydrogen Ion (H^{+})	

Each cation can be combined with any of the anions in the above list. For each combination we will need a separate naming scheme. Note also that in the DVD the polyatomic anions have been grouped according to their charge. Compounds made of polyatomic anions bearing different charges are however named using the same principle.

Section 1.20: Binary Ionic Compounds between Main-group Metals and Nonmetals

A **binary ionic compound** is formed using a metal for the cationic species and a non-metal for the anionic species. Here, we consider only main group metal cations (elements in groups 1, 2 and elements Al, Ga of group 13.

Examples of such compounds are NaCl, MgF_2, KBr, $AlCl_3$, etc…

The rules for naming these compounds are:

1) the metal cation is always named first and the nonmetal anion second

2) the cation is named exactly as the element

3) the anion is named by keeping the root of the element name and adding "**ide**" to the **root**.

   Cl is named **chlorine**, hence Cl^{-} is named **chloride**.

   P is named **phosphorous**, hence P^{-3} is named **phosphide**.

   O is named **oxygen**, hence O^{-2} is named **oxide**.

N is named **nitrogen**, hence N^{-3} is named **nitride**.

S is named **sulfur**, hence S^{-2} is named **sulfide**.

Example 1: Name KBr: K is a metal, Br is a nonmetal. Hence, the above rules apply. Note K is a group 1 metal, hence the cation formed from K is K^{+}. Similarly, Br is a group 17 nonmetal, hence the anion formed from Br is Br^{-}. Combining the two ions forms a neutral compound (1 positive and 1 negative charge). Hence KBr is potassium bromide.

Example 2: Name MgF_2: Mg is a metal, F is a nonmetal. Hence, the above rules apply. Note Mg is a group 2 metal, hence the cation formed from Mg is Mg^{+2}. Similarly, F is a group 17 nonmetal, hence, the anion formed from F is F^{-}. Combining the two ions to form a neutral compound requires 1 Mg^{+2} and 2 F^{-} (2 positive charges and 2 negative charges). Hence, MgF_2 is magnesium fluoride.

Deriving the Chemical Formula from the Name: Conversely, we can derive the chemical formula of a binary ionic compound from its name. Consider calcium chloride. Calcium is named first, as it should since it is a metal. Chloride is the anion of the chlorine atom and is named second. Calcium is a group 2 metal, hence, its cation bears a +2 charge (Ca^{+2}). Chlorine is a group 17 non-metal, hence it bears a charge of -1 (Cl^{-}). Hence, it will take two Cl^{-} ions to balance the +2 charge of Ca^{+2}. Hence, the chemical formula: $CaCl_2$.

Example: To derive the chemical formula of magnesium oxide, we note that magnesium is a group 2 metal, while oxide is the anion of oxygen, a group 16 nonmetal. Hence, the magnesium cation is Mg^{+2}, while the oxide anion is O^{-2}. It takes one oxide anion to balance one magnesium cation. Hence, the chemical formula: MgO.

Section 1.21: Periodic Table: Binary Ionic Compounds between Main Group Metals and Nonmetals

Practice making compounds using group 1, 2 or 13 metals and group 15, 16 or 17 nonmetals. Remember the names and chemical formulas of these compounds.

Section 1.22: Binary Ionic Compounds between Transition Metals and Nonmetals

Many transition metals are capable of forming cations bearing different charges. For example, the element Fe can form Fe^{+2} and Fe^{+3} cations. Hence, the element

Fe can in general form two compounds with a given non metal. For instance, compounds made from Fe^{+2} or Fe^{+3} and Cl^- are $FeCl_2$ and $FeCl_3$, respectively.

Note: it takes two Cl^- to balance one Fe^{+2} and three Cl^- to balance one Fe^{+3}.

The systematic name is derived by naming the cation as the transition metal itself followed (without space) by a Roman numeral in parentheses indicating the charge of the metal cation. For example, Fe^{+2} and Fe^{+3} are named <u>iron(II)</u> and <u>iron(III)</u>, respectively.

The same rule applies to the metals in group 13 and group 14, since they also can form multiple ions. For example: Sn^{+2}, Sn^{+4}, Pb^{+2}, Pb^{+4}.

To complicate matters, there is a different way to name such cations. This method is grounded in the history of chemistry and consists in using different names for each of the possible ions. For instance, in cases where only two common ions are observed, the ion with the lowest charge is named by adding "ous" to the latin root of the element. Consider the following non-exhaustive list of cations that are commonly used in the chemical literature.

Formula	Historical Name	Systematic Name
Cu^+	**cuprous** ion	copper(I)
Cu^{+2}	**cupric** ion	copper(II)
Fe^{+2}	**ferrous** ion	iron(II)
Fe^{+3}	**ferric** ion	iron(III)
Hg_2^{+2} or Hg^+	**mercurous** ion	mercury(I)
Hg^{+2}	**mercuric** ion	mercury(II)
Sn^{+2}	**stannous** ion	tin(II)
Sn^{+4}	**stannic** ion	tin(IV)
Pb^{+2}	**plumbous** ion	lead(II)
Pb^{+4}	**plumbic** ion	lead(IV)

You should use the systematic nomenclature but recognize the old names. The systematic nomenclature for naming binary ionic compounds having a transition metal consists in:

1) naming the transition metal using a Roman numeral in parentheses to characterize the charge of the cation,

2) naming the anion using the root of the element followed by "ide".

For example $FeCl_2$ and $FeCl_3$ are called iron(II) chloride and iron(III) chloride, respectively.

Section 1.23: Periodic Table (Binary Ionic Compounds between Transition Metals and Nonmetals)

Practice naming these compounds and note the various cations exhibited by some of the most common transition metals. In particular note that most of these cations, in contrast with cations from Main Group Metals, do not have the same number of electrons as noble gases. You do not need to memorize all these ions, just know how to name them following the systematic nomenclature.

Sections 1.24 - 1.27: Introduction to Polyatomic Ions

Besides the group of ions formed from metals and nonmetals, there exists a large group of ions called polyatomic ions. These polyatomic ions contain different atoms linked together by covalent bonds. The names of these ions, their chemical formula and net charge must be memorized, as they will be encountered throughout this course.

NH_4^+	ammonium ion	OH^-	hydroxide ion
$C_2H_3O_2^-$	acetate ion	MnO_4^-	permanganate ion
CrO_4^{-2}	chromate ion	$Cr_2O_7^{-2}$	dichromate ion
CN^-	cyanide ion	CO_3^{-2}	carbonate ion
NO_3^-	nitrate ion	NO_2^-	nitrite ion
SO_4^{-2}	sulfate ion	SO_3^{-2}	sulfite ion
PO_4^{-3}	phosphate ion	PO_3^{-3}	phosphite ion

To help you memorize these polyatomic ions, note that the common features shown by the last three sets. The ions, whose names end with "ate" contain one oxygen more than those, whose names end with "ite". Note also that the charge is the same for the "ite" and "ate" anions.

Another important series of polyatomic anions is concerned with the anions made using halogens and oxygen. For instance, consider these formed with Cl and O.

ClO_4^-	perchlorate ion	BrO_4^-	perbromate ion	IO_4^-	periodate ion
ClO_3^-	chlorate ion	BrO_3^-	bromate ion	IO_3^-	iodate ion
ClO_2^-	chlorite ion	BrO_2^-	bromite ion	IO_2^-	iodite ion
ClO^-	hypochlorite ion	BrO^-	hypobromite ion	IO^-	hypoiodite ion

The prefix "per" indicates that these compounds have the most oxygen content (4 oxygen atoms per halogen atom). The prefix "hypo" indicates that these compounds have the least oxygen content (1 oxygen atom per halogen atom). Note again that chlorate, bromate and iodate have one more oxygen atom that chlorite, bromite and iodite.

The final series of polyatomic anions to consider are the series of anions such as:

$$CO_3^{-2}, SO_4^{-2}, SO_3^{-2}, PO_4^{-3}, PO_3^{-3}.$$

Consider the phosphate ion: PO_4^{-3}

Adding H^+ to PO_4^{-3} results in the formation of HPO_4^{-2} or hydrogen phosphate.
Adding H^+ to HPO_4^{-2} leads to the formation of $H_2PO_4^-$ or dihydrogen phosphate.

Similarly, the following compounds are:

HCO_3^-	hydrogen carbonate
HSO_4^-	hydrogen sulfate
HSO_3^-	hydrogen sulfite
HPO_3^{-2}	hydrogen phosphite
$H_2PO_3^-$	dihydrogen phosphite

Sections 1.28 - 1.43: Ionic Compounds Made with Polyatomic Ions

Practice, practice, practice…

Sections 1.44 - 1.45: Introduction to Binary Molecular Compounds

A binary molecular compound is made by combination of two nonmetals. The atoms in a binary molecular compounds are bonded together by covalent bonds. Binary molecular compounds containing hydrogen have the following formulae and names.

Nonmetal	Compound Formula	Compound Name
N	NH_3	Ammonia
O	H_2O	Water
S	H_2S	Hydrogen sulfide
F	HF	Hydrogen fluoride Hydrofluoric acid (in solution)
Cl	HCl	Hydrogen chloride Hydrochloric acid (in solution)
Br	HBr	Hydrogen bromide Hydrobromic acid (in solution)
I	HI	Hydrogen iodide Hydroiodic acid (in solution)

Note that in binary molecular compounds made with hydrogen and nonmetals, hydrogen is usually named first.

Sections 1.46 - 1.47: Introduction to Molecular Compounds Containing Polyatomic Anions and Hydrogen

$HClO_4$	perchloric acid	$HBrO_4$	perbromic acid	HIO_4	periodic acid
$HClO_3$	chloric acid	$HBrO_3$	bromic acid	HIO_3	iodic acid
$HClO_2$	chlorous acid	$HBrO_2$	bromous acid	HIO_2	iodous acid
$HClO$	hypochlorous acid	$HBrO$	hypobromous acid	HIO	hypoiodous acid

H_3PO_4	phosphoric acid
H_3PO_3	phosphorous acid
H_2SO_4	sulfuric acid
H_2SO_3	sulfurous acid
HNO_3	nitric acid
HNO_2	nitrous acid
H_2CO_3	carbonic acid
HCN	hydrogen cyanide
$HC_2H_3O_2$	acetic acid

Note:

1. Acids whose names finish with “ic” come from anions whose names finish in “ate”.
2. Acids whose names finish with “ous” come from anions whose names finish in “ite”.

Examples: Sulfuric acid (H_2SO_4) comes from the sulfate ion (SO_4^{-2}).
Sulfurous acid (H_2SO_3) comes from the sulfite ion (SO_3^{-2}).

Sections 1.50 - 1.51: Concept of Molecule (Part II): Naming Binary Molecular Compounds

The following rules are used when naming binary molecular compounds.

1) The first element is named using its full name (as if the element was a Main Group metal).

2) The second non-metal is named as if it were the anion of a non metal.

3) Greek prefixes are used to indicate the number of atoms present in the compound chemical formula.

4) The prefix “mono” for 1 is never used in front of the name of the first element.

Greek Prefixes:

1	mono	6	hexa
2	di	7	hepta
3	tri	8	octa
4	tetra	9	nona
5	penta	10	deca

Examples:

CO	carbon monoxide
CO_2	carbon dioxide
N_2O_3	dinitrogen trioxide
SO_3	sulfur trioxide
SF_6	sulfur hexafluoride

Exceptions to the Rules for Binary Molecular Compounds:

(1) The oxides of phosphorus are often named assuming phosphorus was a metal. Hence, P_4O_{10} would be phosphorus(V) oxide.

(2) Molecular formula of binary hydrogen compounds with group 15 non-metals are written with the group 15 non-metal first (i.e. NH_3 for ammonia, PH_3 for phosphine).

(3) No Greek prefix is used with the binary hydrogen compounds (for example, H_2S is hydrogen sulfide and NOT dihydrogen monosulfide).