AP Chemistry Preamble

What is Chemistry?

Chemistry can be described as the science that deals with matter, and the changes that matter undergoes. It is sometimes called the *central science* because so many naturally occurring phenomena involve chemistry and chemical change.

Scientific Problem Solving

Scientific (logical) problem solving involves three steps:

1. State the problem and make observations. Observations can be *quantitative* (those involving numbers or measurement) or *qualitative* (those not involving numbers - only language).
2. Formulate a possible explanation (this is known as a *hypothesis*)
3. Perform experiments to test the hypothesis. The results and observations from these experiments lead to the modification of the hypothesis and therefore further experiments.

Eventually, after several experiments, the hypothesis may graduate to become a *theory*. A theory gives a universally accepted explanation of the problem. Of course, theories should be constantly challenged and may be refined as and when new data new scientific evidence comes to light.

Theories are different to *laws*. Laws state what general behavior is observed to occur naturally. For example, the *law of conservation of mass* exists since it has been consistently observed that during all chemical changes mass remains unchanged (i.e., it is neither created or destroyed).

States of Matter and Particle Representations

All matter has two distinct characteristics. It has mass and it occupies space. Properties associated with the three states of matter, and the behaviors of the particles that make up each, are summarized below.

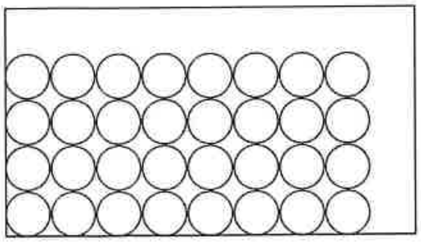
SOLIDS - have a definite shape and definite volume. The particles in a solid are packed tightly together and only vibrate relatively gently around fixed positions.

LIQUIDS - Have no shape of their own but take the shape of their container. A liquid has a definite volume. The particles in a liquid are in free to move around one another.

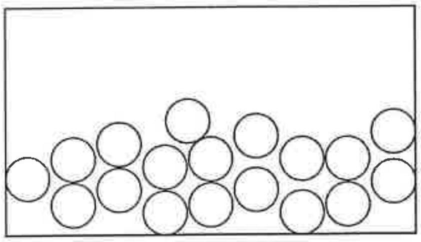
Gases - have neither definite shape nor a definite volume. The particles in a gas spread apart filing all the space of a container available to them and interactions between the particles are considered to be negligible.

Solids, Liquids, and Gases Partical Diagrams

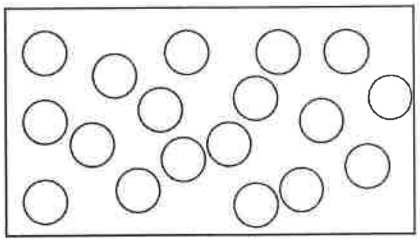
The circles in the diagrams below represent the relative positions and movements of the particles in the three states of matter. Expect to see many such *particulate representations* during the AP course.



Solid



Liquid



Gas

Physical and Chemical Changes and Properties

All matter exhibits physical and chemical properties by which it can be classified. Examples of *physical properties* are color, odor, density, hardness, solubility, melting point, and boiling point.

Chemical properties are those exhibited when a substance reacts with other substances. Examples of chemical properties are reactions with acids and bases, oxidation and reduction (REDOX) and a huge number of other chemical reactions. Changes in which the physical or chemical properties of a substance are altered are considered physical or chemical changes repectively.

Physical Change

If some aspect of the physical state of matter is altered, but the chemical composition remains the same, then the change is considered to be a physical change. The most common physical changes are changes of state. The are summarized below:

SOLID → LIQUID	**Melting**
LIQUID → GAS	**Boiling**
GAS → LIQUID	**Condensing**
SOLID → GAS	**Sublimation**
GAS → SOLID	**Reverse sublimation or deposition**
LIQUID → SOLID	**Freezing**

In solids, the particles have relatively little energy and vibrate around fixed positions. If a solid is heated, the particles gain energy, move around more, and eventually gain enough energy to break away from their fixed positions and form a liquid. Continued heating leads to the liquid particles gaining sufficient energy to break away from one another and form a gas. In a gas the particles move freely and with relatively large amounts of energy.

Chemical Change

In a chemical change, which is often called a chemical reaction, the atoms of a substance are rearranged to form new substances. A chemical change requires that the new substance of substances formed have a different chemical composition to the original substance or substances. Chemical changes are often accompanied by observable changes such as color changes and energy changes.

There is a very important distinction to be made between these two types of change that you will come across later, but for now, note the following:

1. During physical changes, the intermolecular forces (the forces between particles) are disrupted, e.g., boiling water separates one water molecule (H_2O) from another water molecule, but **does not** break any, individual water molecule apart to form hydrogen and oxygen.
2. During chemical changes, the intra forces (the forces *within* substances) are disrupted, e.g., during the electrolysis of water, one water molecule (H_2O) splits up to form O and H atoms. Individual water molecules **do** break apart.

Measurements and Calculations

Measurements, and subsequently calculations applied to those measurements, allow the determination of some of the quantitative properties of a substance e.g., mass, density etc. Measurements are recorded *during* experiments, in real time, but calculations can be performed at any time, and often after the experiment has finished.

Scientific Notation

Measurements and calculations in chemistry often require the use of very large or very small numbers. In order to make handling them easier, such numbers can be expressed using *scientific notation*. All numbers expressed in this manner are represented by a number between 1 and 10 which is then multiplied by 10, raised to a particular power.

The number of places the decimal point moved determines the power of 10. If the decimal point has moved left then the power is positive, if it has moved right then it is negative.

For example, the number 42000.0 is converted fo scientific notation by using the number 4.2 in the process the decimal point has moved four places to the *left*, so the power of 10 used is +4.

$$42000.0 = 4.2 \times 10^4$$

The number 0.00012 is converted to scientific notation by using the number 1.2. In the process the decimal point has moved four places to the *right*, so the power of 10 used is -4.

$$0.00012 = 1.2 \times 10^{-4}$$

Practice - Convert the following to Scientific Notation

a) 24500
b) 356
c) 0.000985
d) .222
e) 12200

Convert the following scientific notation to non-scientific notation numbers

a) 4.2×10^{3}
b) 2.15×10^{-4}
c) 3.14×10^{-6}
d) 9.22×10^{5}
e) 9.57×10^{2}

SI Units

Units tell us the nature of the quantity being measured. Prefixes are used to make writing very large or small numbers easier. Some common SI (*System International*) units and prefixes are given below.

Base quantity	**Name of Unit**	**Symbol**
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s
Amount of substance	Mole	mol
Temperature	Kelvin	K

Prefix	**Symbol**	**Meaning**
Giga	G	10^{9}
Mega	M	10^{6}
Kilo	k	10^{3}
Deci	d	10^{-1}
Centi	c	10^{-2}
Milli	m	10^{-3}
Micro	μ	10^{-6}
Nano	n	10^{-9}
Pico	p	10^{-12}

Dimensional Analysis (the factor label method)

One unit can be converted to another unit by using a conversion factor. Application of the simple formula below will allow the conversion of one unit to another. This method of converting between units is called *dimensional analysis* or the *factor-label method.*

(unit a) (conversion factor) = unit b

The conversion factor is derived from the equivalence statement of the two units. For example, in the equivalence if 1.00 inch = 2,54 cm, the conversion factor will either be,

$$\frac{2.54\ \text{cm}}{1.00\ \text{inch}} \quad \text{or} \quad \frac{1.00\ \text{inch}}{2.54\ \text{cm}}$$

The correct choice is the one that allows the cancellation of the unwanted units. For example, to convert 9.00 inches to cm, perform the following calculation

$$9.00\ \cancel{\text{inches}} \cdot \frac{2.54\ \text{cm}}{1.00\ \cancel{\text{inches}}} = 22.9\ \text{cm}$$

To convert 5.00 cm into inches, perform the following calculation

$$5.00\ \cancel{\text{cm}} \cdot \frac{1.00\ \text{inches}}{2.54\ \cancel{\text{cm}}} = 1.97\ \text{inches}$$

Practice with dimensional analysis

1. Convert the following quantities from one unit to another, using the following equivalence statements; 1.000 m = 1.094 yd, 1.000 mile = 1760 yd, 1.000 kg = 2.205 lbs.
   a. 30m to miles
   b. 1500 yd to miles
   c. 206 miles to m
   d. 34 kg to lbs
   e. 34 lb to kg
2. In each case below, which is the larger quantity?
   a. A distance of 3.00 miles or 3000. M.
   b. A mass of 10.0 kg or 25 lbs.

Temperature

There are two scales of temperature that you will use in AP chemistry. They are Celsius (°C), and Kelvin (K). The following conversion factor is given on the Equations and Constants sheet.

$$K = {}^{\circ}C + 273$$

Temperature Conversion Practice Problems

1. Convert the following temperatures from one unit to the other.
   a. 13 K to °C
   b. 1390 °C to K
2. When discussing a <u>change</u> in temperature, why will it not matter if the change is recorded in Celsius or Kelvin?

Derived Units

All other units can be derived from base quantities. One such unit that is very important is chemistry is volume. Volume has the unit, length ^{3}. Common units for volume are (L) or milliliters (mL).

$$1.000 \text{ mL} = 1.000 \text{ cm}^3$$

And

$$1.000 \text{ L} = 1000 \text{ mL} = 1000. \text{ cm}^3 = 1.000 \text{ dm}^3$$

Density is the ratio of the mass to volume and this equation is found on the Equations and Constants sheet.

$$D = \frac{m}{V}$$

This relationship is particularly useful when dealing with liquids in chemistry. Liquids are most conveniently measured by pouring them into, say, a graduated cylinder. The graduated cylinder records a volume not a mass. In order to calculate the mass of a known volume of a liquid (assuming the density is known) the relationship below can be applied.

Mass = (density) (volume)

Density (con't)

Assuming that density has the units of g/L, volume has units of L, and by using dimensional analysis, it can be seen that the resultant unit for mass in this case is g.

$$\frac{g}{\cancel{L}} \cdot \cancel{L} = g$$

Uncertainty, Significant Figures and Rounding

When reading the scale on a piece of laboratory equipment such as a graduated cylinder or a buret, there is always a degree of uncertainty in the recorded measurement. The reading will often fall between two divisions on the scale and an estimate must be made in order to record the final digit. This estimated final digit is said to be *uncertain* and is reflected in the recording of the numbers by using +/-. All of the digits that can be recorded with certainty are said to be *certain*. The certain and the uncertain numbers taken together are called *significant figures*.

Determining the # of Sig. Fig's present in a

1. Any non-zero integers are always counted as significant figures.
2. Leading zeros are those that precede all non-zero digits and are never counted as significant figures.
3. Captive zeros are those that fall between non-zero digits and are always counted as significant figures.
4. Trailing zeros are those at the end of a number and are only significant if the number is written with a decimal point.
5. Exact numbers have unlimited number of significant figures (Exact numbers are those which are as a result of counting e.g., 3 apples of by definition e.g., 1.000 kg = 2.205 lb).
6. In scientific notation the 10^{x} part of the number is never counted as significant.

Determng the correct #of Sig fig.'s as a result of a Calcuation

1. When multiplying or dividing. Limit the answer to the same number of *significant figures* that appear in the *original* data with the *fewest* number of significant figures.
2. When adding or subtracting. Limit the answer to the same number of *decimal places* that appear in the *original* data with the *fewest* number of *decimal places*.

I.e., don't record a greater degree of significant figures or deciaml places in the calculated answer than the weakest data will allow.

Rounding

Calculators will often present answers to calculations with many more figures than the significant ones. As a result many of the figures shown are meaningless, and the answer, before it is presented, needs to be rounded.

In a multi-step calculation, it is possible to leave the rounding until the end i.e., leave all numbers on the calculator in the intermediate steps, or round to the correct number of figures in each step, or round to an extra figure in each intermediate step and then round to the correct number of significant figures at the end of the calculation. In most cases in the AP chemistry course you will leave numbers on the calculator and round at the end.

Rounding (con't)

Whichever method is being employed, use the simple rule that if the digit directly to the right of the final significant figure is less than 5 then preceding digit stays the same, if it is equal to or greater than 5 the the preseding digit should ne increased by one.

Practice with Sig Fig.'s

1. Determine the number of significant figures in the following numbers.
   a. 250.7
   b. 0.00077
   c. 1024
   d. 4.7×10^{-5}
   e. 34000000
   f. 1003.
2. Use a calculator to carry out the following calculations and record the answer to the correct number of significant figures.
   a. (34.5) (23.46)
   b. 123 / 3
   c. (2.65×10^{-1}) (356)
   d. 21.87 = 45.86
   e. 23.888897 - 11.2
   f. 6 - 3.0

Accuracy & Precision

Accuracy relates to how close the measured value is to the actual value of the quantity. *Precision* refers to how close two or more measurements of the same quantity are to one another.

Practice with Accuracy & Precision

1. Consider three sets of data that have been recorded after measuring a piece of wood that is exactly 6.000 m long.

	Set X	**Set Y**	**Set Z**
	5.864 m	6.002 m	5.872 m
	5.878 m	6.004 m	5.868 m
Average Length	5.871 m	6.003 m	5.870 m

a) Which set of data is the most accurate?

b) Which set of data is the most precise?

Percent Error

The data that are derived in experiments will often differ from the accepted, published, actual value. When this occurs, a common way of expressing accuracy is *percent error*.

$$\% \text{ Error} = \left| \frac{\text{Theoretical Value} - \text{Experimental Value}}{\text{Theoretical Value}} \right| \text{ X } 100$$

AP Chemistry Preamble 2

Atomic Theory

Brief History of Atomic Theory

Circa. 400-5 BC. Greek philosopher Democritus proposes the idea of matter being made up of small, indivisible particles (*atomos*).

Late 18th Century. Lavoisier proposes the Law of conservation of mass and Proust proposes the Law of constant composition.

Early 19th Century. Using the previously unconnected ideas above, John Dalton formulates his Atomic Theory.

Dalton's Atomic Theory

1. **Elements are made from tiny particles called atoms.**
2. **All atoms of a given element are identical (N.B., see isotopes)**
3. **The atoms of a given element are different to those of any other element.**
4. **Atoms of different elements combine to form compounds. A given compound always has the same relative numbers and types of atoms. (Law of constant composition)**
5. **Atoms cannot be created or destroyed in a chemical reaction they are simply rearranged to form a new compounds. (Law of conservation of mass)**

Structure of the atom & the periodic table

Several experiments were being carried out in the 19th & 20th centuries that began to identify the sub-atomic particles that make up the atom. A summary of those experiments is given below.

Scientist	**Experiment**	**Knowledge Gained**	**Relating to**
Crookes	Cathode Ray Tube	Negative particles of some kind exist	Electron
J.J. Thomson	Cathode Ray Deflection	Mass/charge ratio of the electron determined	Electron
Millikan	Oil Drop Experiment	Charge of the electron	Electron
Rutherford, Marsden, & Geiger	Gold Foil Experiment	Nucleus present in atom	The nucleus of an atom and the proton

In the first part of the 20th Century, Bohr built upon Rutherford's idea by introducing quantum theory to the *Solar System Model*, and proposed the idea that the atom was made up of a nucleus containing protons, that was being orbited by electrons, *but only in specific, allowed orbits.* Schödinger quantum model subsequently expanded upon Bohr's model, in order to incorporate the wave nature off the electrons. Once Chadwick's discovered the neutron in 1932, the modern picture of the atom *in its simplest form* was complete.

Particle	**Charge**	**Mass in atomic mass units (amu)**	**Position in atom**
Proton	+1	1	Nucleus
Neutron	0	1	Nucleus
Electron	-1	1/1836	Outside of the nucleus

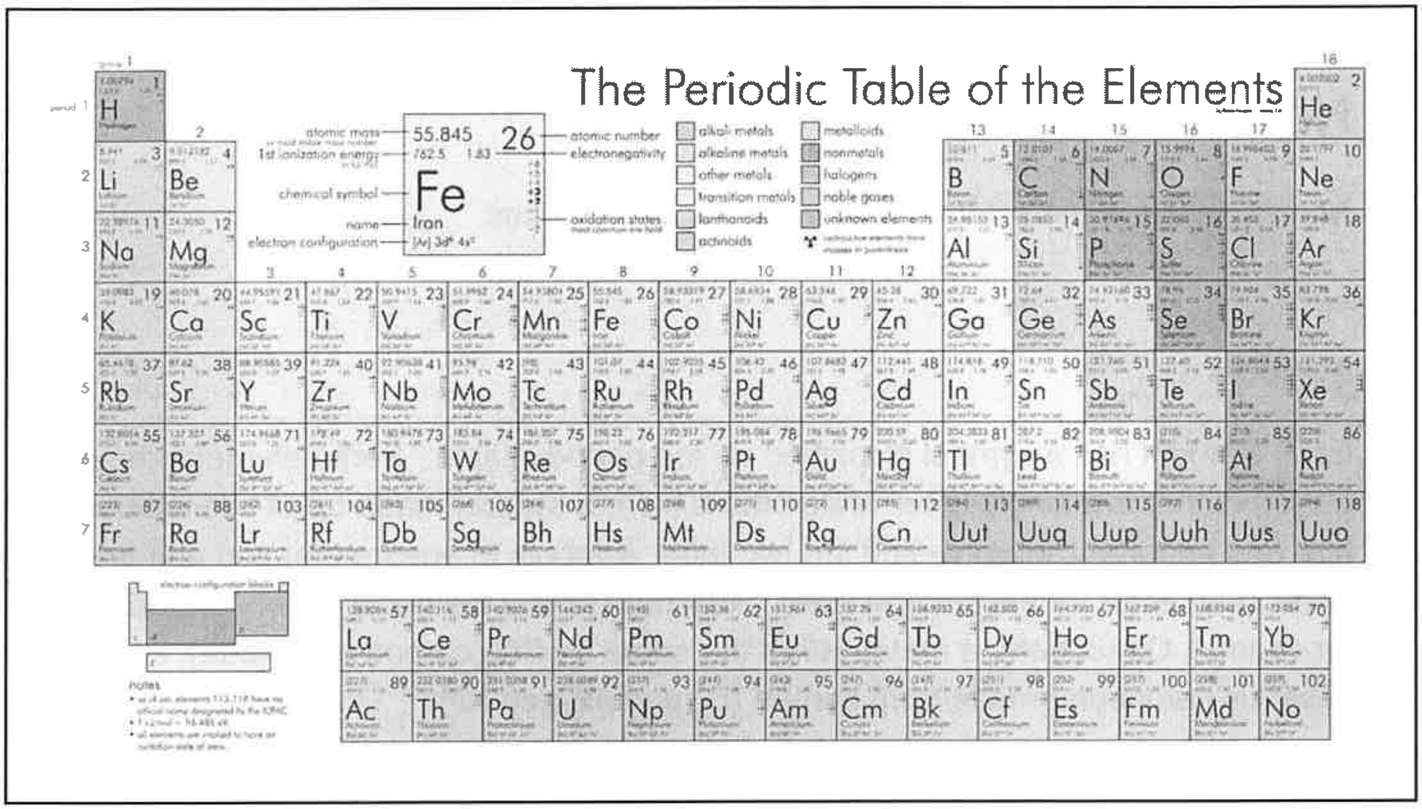
The atomic numbers (in the periodic table below shown above the element symbol and sometimes referred to as Z) and mass numbers (in the periodic table below shown below the symbol and sometimes referred to as A) have specific meanings.

Atomic number = the number of protons in the nucleus of one atom of the element.

Since all atoms are neutral it also tells us the number of electron numbers become unbalanced and the atoms become charged particles, i.e., they are no longer neutral. These charged particles are called *ions*.

A negative ion is formed when an atom gains electrons to possess a greater number of electrons than protons, and is called an *anion*. A positive ion is formed when an atom loses electrons to possess a fewer number of electrons than protons, and is called a *cation*.

Mass number = the number of protons + the number of neutrons in one atom of the element.



Practice Problems

1. Determine the number of protons, electrons, and neutrons in,
   a. $^{210}Pb_{82}$
   b. $^{34}S_{16}$
2. Using only the periodic table above, determine how many elements within the first 20, have atoms with;
   a. The same numbers of protons and electrons
   b. The same number of protons and neutrons

Nomenclature

Nomenclature is the sstem used for naming substances.

Symbols

Each element has a symbol displayed on the periodic table. Some elements have a symbol that is a single letter while others have a symbol made up of two letters. It is very important when writing the two letter symbols to ensure that you use a lower case letter for the second letter. This may sound trivial but is very important, for exampl, Co (cobalt), a metal, is not the same as CO (carbon monoxide), a gaseous compound made from carbon (C) and oxygen (O).

Binary compounds of metals and non-metals (ionic)

Binary compounds are those formed between only two elements. In compounds where one is a metal and one a non-metal an *ionic* compound is formed. An ion is a charged particle and ionic formulae and names can be determined by considering the charge on the ions. To find the formula of an ionic compound the positive and negative charges must be balanced. I. e., there must be no net charge.

To name a binary compound of a metal and an non-metal, the unmodified name of the positive ion is written first followed by the root of the negative ion with the ending modified to -ide. For example, NaCl is sodium chloride.

For a list of ions please refer to the common ion sheet.

Positive ions		**Negative ions**	
Name	Formula	Name	Formula
Hydrogen	H^{+}	Chloride	Cl^{-}
Sodium	Na^{+}	Bromide	Br^{-}
Silver	Ag^{+}	Fluoride	F^{-}
Potassium	K^{+}	Iodide	I^{-}
Lithium	Li^{+}	Hydroxide	OH^{-}
Ammonium	NH_4^{+}	Nitrate	NO_3^{-}
Barium	Ba^{2+}	Oxide	O^{2-}
Calcium	Ca^{2+}	Sulfide	S^{2-}
Copper(II)	Cu^{2+}	Sulfate	SO_4^{2-}
Magnesium	Mg^{2+}	Carbonate	CO_3^{2-}
Zinc	Zn^{2+}		
Lead	Pb^{2+}		
Iron(II)	Fe^{2+}		
Iron(III)	Fe^{3+}		
Aluminium	Al^{3+}		

Most transition metal ions (and a few other metal ions) include a Roam numeral after the name, for example, copper (ii). These metals form ions with varying charges, and the Roman numeral identifies the charge in each case. Elements that commonly form an ion with only a single charge for example, sodium, do not have Roman numerals associated with them.

Practice Nomenclature

1. Name these binary compounds
   a. NaCl
   b. SrO
   c. AlN
   d. $BaCl_2$
   e. K_2O
   f. CuO
   g. Cu_2O

Practice (con't)

2. Convert these names to formulae.

a) magnesium nitride
b) barium bromide
c) aluminum phosphide
d) potassium iodide
e) lithium chloride
f) sodium fluoride
g) tin (IV) bromide

Binary Acids

Acids will be discussed at great length later in the course, but for the purposes of nomenclature, an acid can be defined as a compound that produces hydrogen ions (H^+) when it is dissolved in water, and the formulae of acids start with "H". Binary acids are formed when hydrogen ions combine with monatomic anions.

To name a binary acid use the prefix "hydro" followed by the non-metal name modified to an -ic ending. Then add the word "acid". For example, HCl is hydrochloric acid.

Polyatomic ions

Polyatomic ions are those where more than one element are combined together to create a species with a charge. Some of these ions can be named systematically, others names must be learned. Some common polyatomic ions, their charges and formulae are listed below. For a more complete list check the common ion sheet.

Common Polyatomic Ions

(Memorize!)

Ion	Name
Hg_2^{2+}	mercury (I)
NH_4^+	ammonium
NO_3^-	nitrate
$C_2H_3O_2^-$	acetate
OH^-	hydroxide
CN^-	cyanide
ClO_3^-	chlorate
MnO_4^-	permanganate
NCS^-	thiocyanate
BrO_3^-	bromate

Ion	Name
CO_3^{2-}	carbonate
SO_4^{2-}	sulfate
$Cr_2O_7^{2-}$	dichromate
CrO_4^{2-}	chromate
O_2^{2-}	peroxide
$C_2O_4^{2-}$	oxalate
PO_4^{3-}	phosphate

Polyatomic anions where oxygen is combined with another non-metal are called oxoanions and can be named systematically. In these oxoanions certain non-metals (Cl, N, P, and S) form a series of oxoanions containing different numbers of oxygen atoms. Their names are related to the number of oxygen atoms present, and are based upon the system below.

Name	**Number of oxygen atoms**
Hypo(element)ite	▼
(element)ite	Increase in number of oxygen atoms
(element)ate	
Per(element)ate	▼

Where there are only two members in such a series the endings are - ite and - are. For example, sulfite (SO_3^{2-}) and sulfate (SO_4^{2-}). When there are four members in the series the hypo- and per prefixes are used additionally.

Some oxoanions contain hydrogen and are named accordingly, for example, HPO_4^{2-}, hydrogen phosphate. The prefix thio- means that a sulfur atom has replaced an atom of oxygen in an anion.

To name an ionic compound that contains a polyatomic ion, the unmodified name of the positive ion is written first followed by unmodified name of the negative ion. For example, K_2CO_3 is potassium carbonate.

Oxoacids

Oxoacids are formed when hydrogen ions combine with polyatomic oxoanions. This gives a combination of hydrogen, oxygen and another non-metal.

To name an oxoacid use the name of the oxoanion and replace the -ite ending with -ous or the -ate ending with -ic. Then add the word "acid". For example, H_2SO_4 is sulfuric acid.

To illustrate the names of these oxoanions and oxoacids consider the following example using chlorine as the non-metal.

Formula and name of oxoacid		Formula and name of corresponding oxoanion	
$HClO$	Hypochlorous acid	ClO^-	Hypochlorite
$HClO_2$	Chlorous acid	ClO_2^-	Chlorite
$HClO_3$	Chloric acid	ClO_3^-	Chlorate
$HClO_4$	Perchloric acid	ClO_4^-	Perchlorate

Practice

1. What are the formulae for the following ionic compounds?
   a. Ammonium nitrate
   b. Copper (II) bromide
   c. Copper (I) bromide
   d. Zinc hydrogen sulfate
   e. Sodium perchlorate
   f. Copper (II) iodate

2. Convert the following formulae to names

a) $NaNO_3$
b) $KMnO_4$
c) CaC_2O_4
d) $CuSO_4$
e) Cu_2SO_4
f) KNO_2
g) $LiClO_4$

Binary Compounds of two non-metals (molecular)

If the two elements in a binary compound are non-metals, then the compound is *molecular*.

To name a molecular compound of two non-metals, the unmodified name of the first element is followed by the root of the second element with ending modified to -ide. In order to distinguish between several different compounds with the same elements present use the prefixes mono, di tri, tetra, penta, and hexa to represent one, two, three, four, five and six atoms of the element *respectively*. For example, SO_2 is sulfur dioxide.

Some other examples are given below:

Formula	Name
BCl_3	Boron trichloride
CCl_4	Carbon tetrachloride
CO	Carbon monoxide
CO_2	Carbon dioxide
NO	Nitrogen monoxide
NO_2	Nitrogen dioxide

Note that the prefix mono is only applied to the second element present in such compounds, if the prefix ends with "a" or "o", and the element name begins with "a" or "o", then the final vowel of the prefix is often omitted.

Some compounds have trivial names that have come to supersede their systematic names, for example, H_2O is usually "water", not dihydrogen monoxide.

Practice

1. Write the formula or names for the following molecular compounds.
   a. Dinitrogen tetroxide
   b. Phosphorus pentachloride
   c. Iodine trifluoride
   d. Nitrogen dioxide
   e. Dihydrogen monoxide
2. Convert the following formulae to names.
   a. N_2O_5
   b. PCl_3
   c. SF_6
   d. H_2O
   e. Cl_2O

Hydrates

Hydrates are ionic formula units with water molecules associated with them. The water molecules are incorporated into the solid structure of the ions. Strong heating can generally drive off the water in these salts. Once the water has been removed the salts are said to be anhydrous (without water).

To name a hydrate use the normal name of the ionic compound followed by the term "hydrate" with an appropriate prefix to show the number of water molecules per ionic formula unit. For example, $CuSO_4 \cdot 5\ H_2O$ is copper (II) sulfate pentahydrate.