I. General Chemistry

(1) Scientific Measurement

Scientific Notation

Scientific notation is useful for representing numbers that are very large or numbers that are very small. Scientific notation always begins with a number between 1 and 10, multiplied by ten to an exponent. The exponent is equal to the number of places the decimal was moved from the original number to the position in scientific notation. Positive exponents are used for big numbers (bigger than one) and negative exponents are used for small numbers (smaller than one).

Examples:

The number 450000000 would be represented as 4.5×10^8

Here, the decimal moves 8 places (from after the last zero to between the 4 and the 5) and number is big, so the exponent is positive 8.

The number 0.000362 would be represented as 3.62×10^{-4}

Here, the decimal moves 4 places and the number is small, so the exponent is negative 4.

Metric Measurement

The basic units used in science are the following:

Quantity	**Unit**	**Symbol**
Length	Meter	m
Mass	Gram	g
Volume	Liter	L
Time	Second	s

The Metric System uses prefixes to represent powers of ten:

Prefix	**Symbol**	**Value**
Giga	G	1000000000 10^9
Mega	M	1000000 10^6
kilo	k	1000 10^3
hecto	h	100 10^2
Deca	da	10
Basic Unit	m, g, L, s	1
deci	d	0.1 10^{-1}
centi	c	0.01 10^{-2}
milli	m	0.001 10^{-3}
micro	µ	0.000001 10^{-6}
nano	n	0.000000001 10^{-9}

Metric Conversions

Metric conversions will be done using conversion fractions. Conversion fractions are determined based on metric prefixes. They are equivalent to one, such that when a quantity is multiplied by a conversion fraction, the units are changed, but the value remains unchanged.

ex. 1 **k**g = **1000** g

This equality can be written as two different conversion fractions: $\frac{1kg}{1000\,g}$ or $\frac{1000\,g}{1kg}$

ex. 1 **c**m = **0.01** m

This equality can be writing as two different conversion fractions $\frac{1cm}{0.01m}$ or $\frac{0.01m}{1cm}$

Only one of the conversion fractions will be used to convert from the units of the given quantity to the new units. The conversion fraction is chosen such that when it is multiplied by the given quantity, the units of the given quantity will cancel and the new units will remain. Generally this will mean the units of the given quantity will be in the denominator of the conversion fraction and the new units will be in the numerator of the conversion fraction.

Conversion to/from a Base Unit
ex. Convert 2.5 kg to g

2.5 kg x $\frac{1000\,g}{1\,kg}$ = 2500 g

(this conversion fraction is used so that the units of kg cancel and the units of g remain, 2.5 is multiplied by 1000 to give 2500)

ex. Convert 0.50 m to cm

0.50 m x $\frac{1\,cm}{0.01\,m}$ = 50 cm

(this conversion fraction is used so that the units of m cancel and the units of cm remain, 0.50 is divided by 0.01 to give 50)

Conversions between Two Units
Conversions between two units are done in two steps:
(1) The given unit must be first converted to the base unit (g, m ,L, or s).
(2) The answer in the base unit is then converted to the new unit.

ex. Convert 80.0 cm to km

(1) Convert cm to m

80.0 cm x $\frac{0.01\,m}{1\,cm}$ = 0.800 m

(2) Convert the answer in m to km

0.800 m x $\frac{1\,km}{1000\,m}$ = 0.000800 km

8.00×10^{-4} km

These two steps can be combined into one calculation:

80.0 cm x $\frac{0.01\,m}{1\,cm}$ x $\frac{1\,km}{1000\,m}$ = 0.000800 km or 8.00×10^{-4} km

ex. Convert 0.44 Mg to dag

$$0.44\,Mg\; x\, \frac{1x10^{6}\,g}{1\,Mg}\, x\, \frac{1\,dag}{10\,g} = 4.4x10^{4}\,dag$$

Conversions with Squared and Cubed Units
To convert from one squared unit to another or from one cubed unit to another, the conversion fraction must be squared in order for the units to cancel. It is important to remember to square/cube any numbers in the conversion factor as well.

ex. 200. cm^2 = ____________ m^2

X WRONG:

200. cm^2 x $\frac{0.01\,m}{1\,cm}$ = 2 .00 m^2 → here the units DO NOT CANCEL so the work/answer is incorrect.

√ RIGHT:

200. cm^2 x $\left(\frac{0.01\,m}{1\,cm}\right)^2$ = 200 cm^2 x $\frac{0.0001\,m^2}{1\,cm^2}$ = 0.200 m^2 → here the units DO CANCEL so the work/answer is correct.

General Conversions
Conversion factors can be applied to conversions between any units.

ex . (a) A horse weighs 1100. pounds, what is the mass in kilograms? (1 kilogram = 2.2046 pounds)
(b) How many tons does a horse weight? (1 ton = 2000 pounds)
(c) How many metric tons does a horse weigh (1 metric ton = 1000 kilograms)

(a) $1100.\,pounds\ x\ \dfrac{1\,kg}{2.2046\,pounds} = 499.0\,kg$

(b) $1100.\,pounds\ x\ \dfrac{1\,ton}{2000\,pounds} = 0.5500\,tons$

(c) $499.0\,kg\ x\ \dfrac{1\,metric\,ton}{1000\,kg} = 0.4990\,metric\,tons$

ex. A playing card measures 9.0 cm by 6.0 cm, what is the area in centimeters squared? in inches squared? (1 inch = 2.54 centimeters)

$$A = l\ x\ w = 9.0\,cm\ x\ 6.0\,cm = 54\,cm^2$$

$$54\,cm^2\ x \left(\frac{1\,inch}{2.54\,cm}\right)^2 = 54\,cm^2\ x\ \frac{1\,inch^2}{6.4516\,cm^2} = 8.4\,inch^2$$

ex. The speed limit reads 90 km/h, what is the speed in miles/h? in m/s? (1 mile = 1.6093 kilometers)

$$\frac{90\,km}{hour}\ x\ \frac{1\,mile}{1.6093\,km} = 56\,miles/hour \qquad \frac{90\,km}{hour}\ x\ \frac{1000\,m}{1\,km}\ x\ \frac{1\,hour}{3600\,s} = 25\,m/s$$

ex. A car has a fuel rating of 35.0 miles per gallon. What is the rating in kilometers per liter? (1 gallon = 3.7854 liters)

$$\frac{35.0\,miles}{Gallon}\ x\ \frac{1.6093\,km}{1\,mile}\ x\ \frac{1\,Gallon}{3.7854\,L} = 14.9\,km/L$$

Temperature

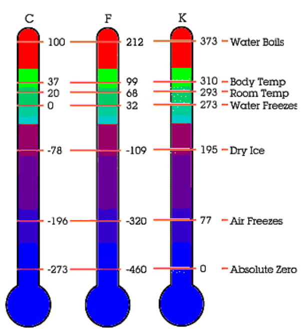
There are three scales for measuring temperature.
(1) **Fahrenheit**, Gabriel Fahrenheit (1686–1736)
0 F was set as the coldest temperature that could be obtained in a liquid in the lab, a mixture of salt, ice, and water; 32 F was set as the temperature of a mixture of ice and water.
(2) **Celsius**, Anders Celsius (1701–1744)
0 ºC was set as the freezing point of water and 100 ºC was set as the boiling point of water.
(3) **Kelvin**, Lord Kelvin, William Thomson (1824 –1907)
0 K (also called "absolute zero") is the coldest temperature that is possible. At 0 K all molecular motion stops.

Temperature Conversions
The following equations can be used to convert between the different temperature scales.

F = $\dfrac{9}{5}(C) + 32$ C = $\dfrac{5}{9}(F - 32)$ K = C + 273

F = Fahrenheit
C = Celsius
K = Kelvin

Some important temperatures represented in each of the temperature scales.



Temperature Conversions
Complete the following temperature conversions.

ex. _____ °C = 86 F = _____ K

$$C = \frac{5}{9}(F - 32)$$
$$= \frac{5}{9}(86 - 32)$$
$$= 30\ °C$$

K = C + 273
= 30 + 273
= 303 K

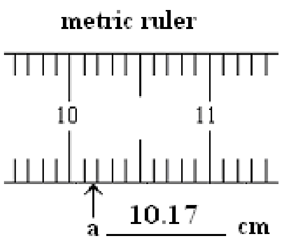
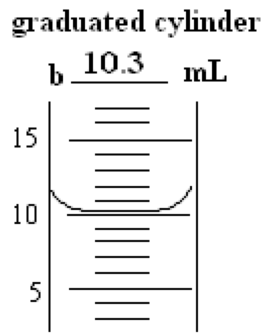
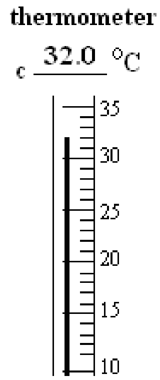
ex. _____ °C = _____ F = 318 K

K = C + 273
318 = C + 273
C = 45 °C

$$F = \frac{9}{5}(C) + 32$$
$$= \frac{9}{5}(45) + 32$$
$$= 113\ F$$

Reading Metric Scales
Metric scales (i.e. rulers, thermometers, graduated cylinders, etc) can only be read to a certain accuracy determined by the size of the smallest division. The rule to follow is that the scale can be read to <u>one place past the smallest division</u>.

Examples:

		
Metric ruler: for *this* ruler, the smallest division is 0.1 cm, therefore the ruler can be read to the nearest 0.01 cm (for *this* ruler, the reading must have TWO decimal places).	Graduated cylinder: for *this* graduated cylinder, the smallest division is 1 mL, therefore, the graduated cylinder can be read to the nearest 0.1 mL (for *this* graduated cylinder, the reading must have ONE decimal place).	Thermometer: for *this* thermometer, the smallest division is 1 °C, therefore the thermometer can be read to the nearest 0.1 °C (for *this* thermometer, the reading must have ONE decimal place).

The digits in a reported reading are called **significant figures**. The number of significant figures reported in collected data is very important as it provides an indication of the level of precision in the experiment. There are very specific rules used in science for counting and performing mathematical operations with significant figures. All calculations that are performed will be rounded to an appropriate number of significant figures as determined by the rules and procedures outlined below.

Significant Figures

I. Rules for Counting Significant Figures (significant figures are underlined)
(1) All non–zero digits are significant
ex. <u>55</u> has <u>2</u> sig figsex. <u>123456</u> has <u>6</u> sig sigs

(2) Zeros between non–zero digits are significant
ex. <u>505</u> has <u>3</u> sig figs ex. <u>10006</u> has <u>5</u> sig figs

(3) Zeros before a decimal place are significant
ex. <u>500</u> may have <u>1</u> *or* <u>3</u> sig figs.
To clarify the number of significant figures the number can be written in two possible ways:
1. Use scientific notation (i.e. to write with 1 sig fig use $5x10^2$; to write with 3 sig figs use $5.00x10^2$)
2. Include decimal point after the number to show that all zeros are significant (i.e. to write with 3 sig figs use 500.)

(4) Trailing zeros after a decimal place, after a number are significant
ex. <u>50.0</u> has <u>3</u> sig figs ex. <u>1.000</u> has <u>4</u> sig figs

(5) Leading zeros before/after a decimal place, before a number are NOT significant
ex. 0.0<u>55</u> has <u>2</u> sig figs ex. 0.<u>140</u> has <u>3</u> sig figs

(6) Powers of ten in scientific notation are not significant
ex. <u>5.5</u>$x10^3$ has <u>2</u> sig figs ex. <u>1.20</u>$x10^5$ has <u>3</u> sig figs

Rounding
When rounding to a given place value, the only digit that needs to be considered is the number directly following the digit being rounded. If the number following the digit being rounded is 5 or greater, the value is rounded up by one. If the number following the digit being rounded is four or less, the value is rounded down and stays as is.
Round 1248.5397 to the nearest….

thousand	hundreds	tens	ones	tenth	hundredth	thousandth
= 1000 can also be written as $1x10^3$	= 1200 can also be written as $1.2x10^3$	= 1250 can also be written as $1.25x10^3$	= 1249 can also be written as $1.249x10^3$	= 1248.5 can also be written as $1.2485x10^3$	= 1248.54 can also be written as $1.24854x10^3$	= 1248.540 can also be written as $1.248540x10^3$
here, the number being rounded, 1, is followed by a 2, so the 1 stays	here, the number being rounded, 2, is followed by a 4, so the 1 stays	here, the number being rounded, 4, is followed by an 8, so 4 rounds up to 5	here, the number being rounded, 8, is followed by a 5, so 8 rounds up to 9	here, the number being rounded, 5, is followed by a 3, so the 5 stays	here, the number being rounded, 3, is followed by a 9, so 3 rounds up to 4	here, the number being rounded, 9, is followed by a 7, so 9 rounds up to 10 (and therefore 3 before changes to 4)

ex. Round each value to the indicated number of significant figures

(a) 5.589 (2)
= 5.6

(b) 0.012561 (3)
= 0.0126

(c) 11.213 (2)
= 11

(d) $6.72x10^3$ (2)
= $6.7x10^3$

(e) 0.134999 (2)
=0.13

(f) 25601 (2)
= $2.6x10^4$

II. Operations with Significant Figures

(1) Addition/Subtraction
– the answer is rounded to the fewest number of *decimal places*

ex. 26.46 + 4.123 = 30.583 → 30.58 (rounded to 2 decimal places)

ex. 100.5 – 75.416 = 25.084 → 25.1 (rounded to 1 decimal place)

(2) Multiplication/Division
– the answer is rounded to the fewest number of *significant figures*

ex. 5.83 x 2.3 = 13.409 → 13 (rounded to 2 significant figures)

ex. 100.5 ÷ 20.2 = 4.9752… → 4.98 (rounded to 3 significant figures)

(2) Density
Density: the mass of a given volume of matter.

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

D = density
m = mass
V = volume

ex. An empty beaker has a mass of 200.0 g. When some ethanol (density 0.800 g/mL) is poured in the beaker, the mass of the ethanol and the beaker is 250.0 g. What is the volume of ethanol in the beaker? What is the volume in L?

mass = 250.0 g – 200.0 g = 50.0 g

$$V = \frac{m}{D} = \frac{(50.0\,g)}{0.800\,g/mL} = 62.5\,mL$$

$$62.5\,mL\ x\ \frac{1L}{1000\,mL} = 0.0625\,L$$

ex. A copper block has a mass of 50.0 g. When the block is placed in a graduated cylinder with 25.00 mL of water, the water level rises to 30.59 mL. What is the density of copper? What is the density in kg/L?

V = 30.59 mL – 25.00 mL = 5.59 mL

$$D = \frac{m}{V} = \frac{50.0g}{5.59\,mL} = 8.94\,g/mL$$

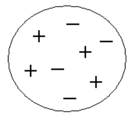
$$\frac{8.94\,g}{mL}\ x\ \frac{1kg}{1000\,g}\ x\ \frac{1000\,mL}{1L} = 8.94\,kg/L$$

ex. A block of lead (density = 11.34 g/cm^3) has sides of length 0.50 cm, 2.0 cm. and 0.50 cm. What is the mass? What is the mass in kg?

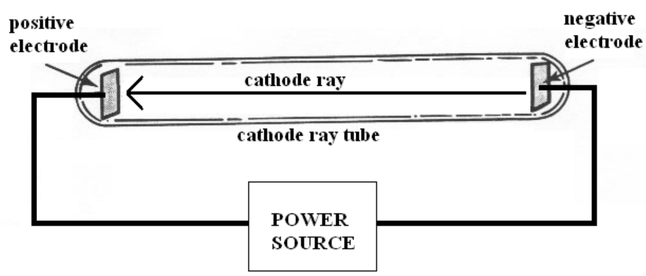
$$V = l\,x\,w\,x\,h = (0.50\,cm)(2.0\,cm)(0.50\,cm) = 0.50\,cm^3$$
$$m = DxV = (11.34\,g/cm^3)(0.50\,cm^3) = 5.7\,g$$
$$5.7\,g\,x\,\frac{1\,kg}{1000\,g} = 0.057\,kg$$

(3) Atomic Structure
John Dalton (1766–1844): Atomic Theory
Atomic Theory
(1) All matter is made of atoms.
(2) Each element has its own kind of atom.
(3) Atoms cannot be created or destroyed.
(4) Molecules are formed when atoms are combined (compounds are formed when elements are combined).

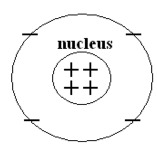
Joseph John Thomson (1856–1940): discovered the **electron**
"Plum Pudding Model"
An atom has lightweight negative particles (electrons) scattered throughout a heavy positive matrix (protons).



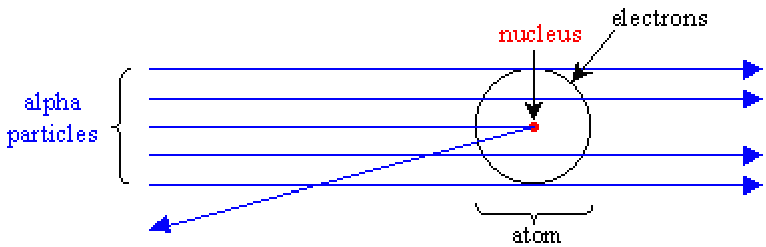
Thomson's Cathode Ray Experiment
When a high voltage power supply was applied to a tube containing a very small amount of gas, called a cathode ray tube, a ray of particles (i.e. a cathode ray) was produced. The cathode ray originated from the negative electrode of the tube, and was repelled away from the negative electrode and attracted towards the positive electrode. Thomson proposed that the cathode ray was composed of negatively charged particles from the atoms of the metal at the electrode. These negatively charged particles were later called electrons. Since it was known that atoms are neutral (uncharged), Thomson deduced that atoms must also contain positively charged particles. These positively charged particles were later called protons.



Ernest Rutherford (1871–1937): discovered the **nucleus**
"Nuclear Model"
An atom has a small, positively charged nucleus containing protons (which makes up the majority of the mass) surrounded by a cloud of negatively charged electrons (which take up most of the space).



Rutherford's Gold Foil Experiment
A layer of gold foil was bombarded by fast moving alpha particles (which are positively charged). It was expected that the alpha particles would pass through the gold foil. What was observed was that *most* alpha particles passed through the foil, but some of the alpha particles (about 1 in 8000) were actually sent backwards once they hit the foil. Rutherford proposed that the particles sent backwards had hit a small, positively charged "nucleus".

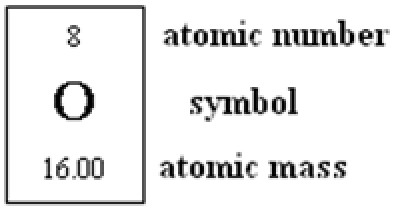


James Chadwick (1891–1974): discovered the **neutron**, a neutral particle in the nucleus which contributes to the mass of an atom.

Properties of Subatomic Particles

Subatomic Particle	**Charge**	**Relative Mass**	**Location**
Proton	+1	1	nucleus
Electron	–1	1/1836 (approximately zero)	orbiting nucleus
Neutron	0	1	nucleus

Atomic Configuration



Name
Each element has a specific name. New elements are generally named by the discoverer(s), but the names must be approved for adoption by IUPAC (**International Union of Pure and Applied Chemistry**).

Symbol
The chemical symbol is either one UPPERCASE letter or one UPPERCASE letter followed by one lowercase letter.
The symbol may derive from the name of the element in a variety of languages. Some elements have been known since ancient times (for example: copper, gold, lead, silver, iron, tin, sulphur, mercury) others have been discovered much more recently (for example: copernicium: 1996, flerovium: 1998, and livermorium: 2000).

Atomic Number
The atomic number is different for each element. The atomic number is always equal to the number of protons. For a neutral (uncharged) atom, the atomic number is also equal to the number of electrons.

Atomic Mass
The atomic mass can be found below an element on the periodic table. The atomic mass is the combined mass of all the subatomic particles. Since electrons do not significantly contribute mass, that atomic mass is sum of the number of protons and the number of neutrons.

$$atomic\ mass = number\ of\ protons + number\ of\ neutrons$$

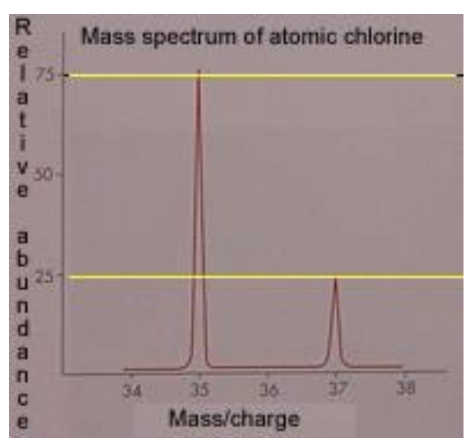
Isotopes
Isotopes: Different "versions" of atoms of the same element that have the same number of protons but different numbers of neutrons and therefore different atomic masses.

The atomic mass listed on the periodic table is actually the (weighted) average atomic mass of all of the isotopes.
The average atomic mass is calculated using the following equation:

$$Average\ Atomic\ Mass = (mass)(abundance) + (mass)(abundance)\ldots$$

The number of terms in the equation will depend on the number of isotopes an element has.
The abundance of an isotope is given as a percentage and indicates how common the isotope is.
The name of an isotope is written by giving the name of an element followed by its mass.

Scientists use an instrument called a mass spectrometer to determine the exact mass and abundance of the different isotopes of elements. The mass spectrum for chlorine is shown.



ex. There are two isotopes of chlorine: chlorine-35 and chlorine-37. Both isotopes of chlorine have 17 protons because all atoms of chlorine have 17 protons. Chlorine-35 has a mass of 35 and 18 neutrons (35 – 17 = 18). Chlorine-37 has a mass of 37 and 20 neutrons (37 – 17 = 20).

Chlorine–35 is found to have an exact mass of 34.968853 a.m.u. and an abundance of 75.77%, chlorine–37 is found to have an exact mass of 36.965903 and an abundance of 24.23%. (This means is there were 10000 atoms of chlorine, 7577 would be chlorine–35 and 2423 would be chlorine–37. Abundances always add up to 100%) The abundances and the masses can be used to show how the average atomic mass on the periodic table is calculated.

$$Average\ Atomic\ Mass = (34.968853)(0.7577) + (36.965903)(0.2423)$$
$$= 35.45\ a.m.u.(as\ given\ on\ the\ periodic\ table)$$

In the equation, the percentages are written as decimals. To turn a percent into a decimal divide by 100. 75.77/100 = 0.7577 and 24.23/100 = 0.2423

Nuclide Symbols

A nuclide symbol gives the mass and the number of protons for a given isotope of an atom.

ex. The nuclide symbol for chlorine-35 $^{35}_{17}Cl$ is and the nuclide symbol for chlorine-37 is $^{37}_{17}Cl$

Ions: Charged atoms.

Atoms are most stable when they have a full valence shell (like the *noble gases*). Atoms can lose or gain electrons to establish a full valence shell. When electrons are lost, the atom becomes a positively charged **cation**. When electrons are gained, the atom becomes a negatively charged **anion**.

ex. Na (11 protons, 11 electrons) → Na^+ (11 protons, 10 electrons) has the same electronic structure as neon
ex. Cl (17 protons, 17 electrons) → Cl^- (17 protons, 18 electrons) has the same electronic structure as argon

(4) Periodic Table

Classification of Elements

Elements on the periodic table can be classified as metal, non–metals, or metalloids

Metal	**Non–Metal**	**Metalloid**
- elements to the left of the "staircase" - malleable (can be flattened into sheets) - ductile (can be pulled into wires) - shiny - good conductors of heat and electricity (conductors)	- elements to the right of the "staircase" - brittle - dull - poor conductors of heat and electricity (insulators)	- found along the "staircase" - include B, Si, Ge, As, Sb, and Te - have some properties of metals (i.e. most are good at conducting electricity) and some properties of non-metals (i.e. most are brittle) - most are semi-conductors which means that their conductivity increases at higher temperature (whereas the conductivity of a metal decreases at higher temperatures).

Elements in the periodic table are arranged into **groups/families** (vertical) and **rows/period**s (horizontal)
Many groups of elements on the periodic table have similar properties and are called "chemical families".

Group I/1A: Alkali Metals (Li, Na, K, Rb, Cs, Fr)
- soft, shiny, silver metals
- have one valence electron; form 1+ ion
- very reactive

Group II/2A: Alkaline Earth Metals (Be, Mg, Ca, Sr, Ba, Ra)
- shiny, silvery-white metals
- have two valence electrons; form 2+ ion
- fairly reactive

Group 3-12: Transition Metals
- most are silver metals (except copper and gold)
- most metals can display a variety of charges when forming ions (although some metals are generally found with only one common charge: Silver: Ag^+, Zinc: Zn^{2+}, Cadmium: Cd^{2+}, and Scandium: Sc^{3+})
- many ions display a variety of colours in solution, for example copper (II), Cu^{2+} is blue and nickel (II), Ni^{2+} is green, notable exceptions are silver Ag^+ and zinc Zn^{2+}, both of which are have colourless ions

Group VII/7A/17: Halogens (F, Cl, Br, I, At)
- F_2 and Cl_2 are gases, Br_2 is a liquid, and I_2 is a solid (At is very rare)
- seven valence electrons; form –1 ion
- very reactive

Group VIII/8A/18: Noble Gases (He, Ne, Ar, Kr, Xe, Rn)
- gaseous
- have a full valence shell so they do not form ions
- very stable and unreactive/inert

(5) Chemical Compounds
Ionic Compounds
- Ionic compounds form between a metal and a non-metal.
- Electrons are transferred from the metal to the non-metal to form ions (in order for each element to have a full valence shell)
- The metal loses electrons to become a positively charged cation and the non-metal gains electrons to become a negatively charged anion.
- Ions are held together by electrostatic attraction of opposite charges (i.e. "opposites attract").

Naming and Writing Chemical Formulas for Ionic Compounds
There are several types of ionic compounds that depend on the types on ions that form the compound.

Simple (Binary) Ionic Compounds
I. Naming Simple Ionic Compounds
(1) Name the metal.
(2) Name the non–metal and change the ending to "ide".

ex. KI
potassium iodide

ex. BaO
barium oxide

ex. MgF_2
magnesium fluoride

ex. $ScCl_3$
scandium chloride

ex. Ga_2O_3
gallium oxide

II. Writing formulas for Simple Ionic Compounds
(1) Write the symbol and charge for the metal.
(2) Write the symbol and charge for the non–metal.
The charge of each ion can be predicted based in the number of valence electrons for the atom/group number on periodic table.
(3) Use the charges to determine the number of atoms of each element required: the total positive charge must be equal to the total negative charge.
(4) Write the formula indicating the number of atoms of each element with subscripts.

ex. lithium bromide = Li^+ Br^- = LiBr
The charges balance requiring only one atom of each element (a subscript of one is not required).

ex. sodium oxide = Na^+ O^{2-} = Na_2O
Two sodium atoms are needed to give a combined charge of +2 to balance the charge of –2 from one oxygen atom. The formula is written with a subscript of two for sodium.

ex. strontium sulphide = Sr^{2+} S^{2-} = SrS

ex. aluminum chloride = Al^{3+} Cl^- = $AlCl_3$

ex. calcium nitride = Ca^{2+} N^{3-} = Ca_3N_2

Multivalent Ionic Compounds
Multivalent ionic compounds contain a multivalent metal.
Multivalent metal: a metal that can form ions with a variety of charges.
Includes most transition metals, many lanthanides and actinides, as well as other metals such as lead (Pb^{2+}/Pb^{4+}) and tin Sn^{2+}/Sn^{4+}).
Commonly, silver (Ag^+), zinc (Zn^{2+}), cadmium (Cd^{2+}), and scandium (Sc^{3+}) are transition metals that generally display only one charge and do not require a roman numeral when giving the name.

I. Naming Multivalent Metals
(1) Name the metal indicating the charge with a roman numeral. (Use the number of atoms of each ion and the charge of the non–metal to determine the charge and therefore the roman numeral of the metal)

Charge	Roman Numeral
+1	I
+2	II
+3	III
+4	IV
+5	V
+6	VI
+7	VII

(2) Name the non–metal and change the ending to "ide".

ex. CuCl = copper (I) chloride

There is one chlorine atom and chlorine has a charge of –1. The copper must therefore have a charge of +1 and is written with a roman numeral of "I".

ex. $CuCl_2$ = copper (II) chloride

There are two chlorine atoms and each chlorine has a charge of –1, so together they have a total charge of –2. The copper must therefore have a charge of +2 and is written with a roman numeral of II".

ex. FeO = iron (II) oxide

ex. Fe_2O_3 = iron (III) oxide

ex. PbS = lead (II) sulphide

ex. PbS_2 = lead (IV) sulphide

II. Writing formulas for Multivalent Metals
(1) Write the symbol and charge for the metal (**use the Roman Numeral given in the name to determine the charge**).
(2) Write the symbol and charge for the non–metal.
(3) Use the charges to determine the number of atoms of each element required .
(4) Write the formula indicating the number of atoms of each element with subscripts.

ex. gold (I) bromide = Au^{+} Br^{-} = AuBr

The roman numeral given for gold is "I", therefore the charge is +1. The charge for bromine is –1. The charges balance requiring only one atom of each element

ex. gold (III) bromide = Au^{3+} Br^{-} = $AuBr_3$

The roman numeral given for gold is "III", therefore the charge is +3. The charge for bromine is –1. Three bromine atoms are needed to give a combined charge of –3 to balance the charge of +3 from gold. The formula is written with a subscript of 3 for bromine.

ex. nickel (II) sulphide = Ni^{2+} S^{2-} = NiS

ex. nickel (III) sulphide = Ni^{3+} S^{2-} = Ni_2S_3

ex. rhenium (IV) oxide = Re^{4+} O^{2-} = ReO_2

ex. rhenium (VII) oxide = Re^{7+} O^{2-} = Re_2O_7

Polyatomic Ionic Compounds
Polyatomic ionic compounds contain a polyatomic ion.
Polyatomic Ions: groups of elements that combine together to form a single ion. Polyatomic ions are shown below.

+1 Charge		**–2 Charge**	
ammonium	NH_4^{+}	hydrogen phosphate	HPO_4^{2-}
		hydrogen phosphite	HPO_3^{2-}
–1 Charge		carbonate	CO_3^{2-}
dihydrogen phosphate	$H_2PO_4^{-}$	carbonite	CO_2^{2-}
dihydrogen phosphite	$H_2PO_3^{-}$	sulphate	SO_4^{2-}
hydrogen carbonate	HCO_3^{-}	sulphite	SO_3^{2-}
hydrogen carbonite	HCO_2^{-}	chromate	CrO_4^{2-}
hydrogen sulphate	HSO_4^{-}	dichromate	$Cr_2O_7^{2-}$
hydrogen sulphite	HSO_3^{-}	oxalate	$C_2O_4^{2-}$
nitrate	NO_3^{-}	thiosulphate	$S_2O_3^{2-}$
nitrite	NO_2^{-}	silicate	SiO_3^{2-}
hydroxide	OH^{-}		
cyanide	CN^{-}	**–3 Charge**	
cyanate	OCN^{-}	phosphate	PO_4^{3-}
thiocyanate	SCN^{-}	phosphite	PO_3^{3-}
permanganate	MnO_4^{-}	arsenate	AsO_4^{3-}
chlorate	ClO_3^{-}	arsenite	AsO_3^{3-}
chlorite	ClO_2^{-}	borate	BO_3^{3-}
hypochlorite	ClO^{-}		
perchlorate	ClO_4^{-}		
bromate	BrO_3^{-}		
bromite	BrO_2^{-}		
hypobromite	BrO^{-}		
perbromate	BrO_4^{-}		
iodate	IO_3^{-}		
iodite	IO_2^{-}		
hypoiodite	IO^{-}		
periodate	IO_4^{-}		
acetate	$C_2H_3O_2^{-}$ or CH_3COO^{-}		

Ion Colours
Some polyatomic ions also display colours, for example, permanganate, MnO_4^{-}, is purple, chromate, CrO_4^{2-}, is yellow, and dichromate, $Cr_2O_7^{2-}$, is orange

I. Naming Polyatomic Ionic Compounds
(1) Name the metal first (if the metal is multivalent, a Roman Numeral must be included to indicate the charge)
(2) Name the polyatomic ion (note: the ending of the polyatomic does not change).

ex. $CaCO_3$ = calcium carbonate

ex. $NaNO_2$ = sodium nitrite

ex. $Mg(OH)_2$ = magnesium hydroxide

ex. NH_4Cl = ammonium chloride

ex. $CuSO_4$ = copper (II) sulphate

ex. $Hg(IO_3)_2$ = mercury (II) iodate

II. Writing formulas for Polyatomic Ionic Compounds
(1) Write the symbol and charge for the metal.
(2) Write the symbol and charge for the polyatomic ion. (Polyatomic ions can often be recognized from the "ate" or "ite" ending in the name. Commonly, only ammonium, hydroxide, and cyanide do not have these endings.)
(3) Use the charges to determine the number of each ion required.
(4) Write the formula indicating the number of atoms of each element with subscripts (if more than one of a polyatomic ion is required, the formula must be written using **parentheses**)

ex. sodium carbonate = Na^+ CO_3^{2-} = Na_2CO_3

Sodium has a charge of +1. Carbonate has a formula of CO_3^{2-}. where the charge is –2. Two sodium atoms are needed to give a combined charge of +2 to balance the charge of –2 from one carbonate ion. The formula is written with a subscript of two for sodium.

ex. magnesium chlorite = Mg^{2+} ClO_3^- = $Mg(ClO_3)_2$

Magnesium has a charge of +2. Chlorite has a formula of ClO_3^-. where the charge is –1. Two chlorite ions are needed to give a combined charge of –2 to balance the charge of +2 from one magnesium ion. The formula is written with a subscript of two for chlorite, with the chlorite ion written in parenthesis.

ex. ammonium sulphate = NH_4^+ SO_4^{2-} = $(NH_4)_2SO_4$

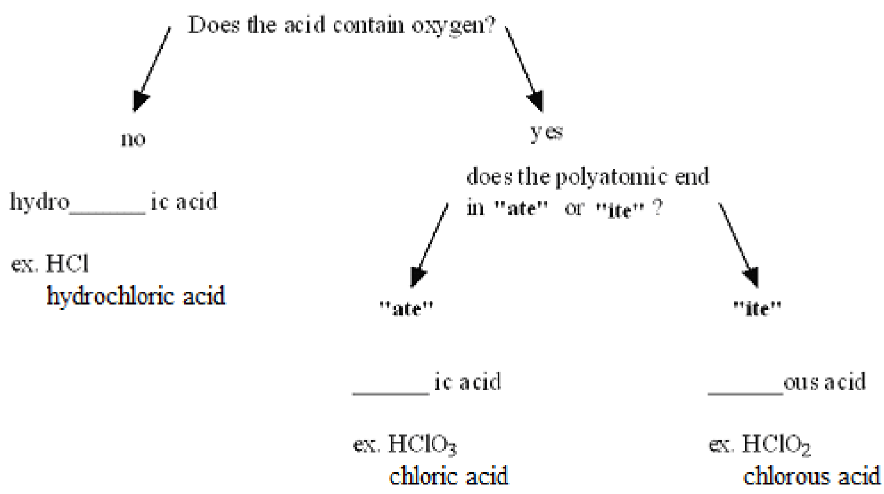
ex. iron (III) phosphate = Fe^{3+} PO_4^{3-} = $FePO_4$

ex. cobalt (II) acetate = Co^{2+} $C_2H_3O_2^-$ = $Co(C_2H_3O_2)_2$

ex. calcium hydrogen carbonate = Ca^{2+} HCO_3^- = $Ca(HCO_3)_2$

Acids
The formula for acids begin with a hydrogen (H^+ ion).

I. Naming Acids
To name an acid, follow the given flow chart:



Practice Examples

(1) HF = hydrofluoric acid	(2) H_3PO_4 = phosphoric acid	(3) HBrO = hypobromous acid

II. Writing formulas for acids
(1) Write the symbol and charge for hydrogen (H^+).
(2) Use the prefix/ending of the name to determine the second part of the compound.
(3) Use the charges to determine the number of each ion required.
(4) Write the formula indicating the number of hydrogen ions with a subscript.

ex. hydrosulphuric acid = H^+ S^{2-} = H_2S ex. sulphuric acid = H^+ SO_4^{2-} = H_2SO_4 ex. sulphurous acid = H^+ SO_3^{2-} = H_2SO_3

Practice Examles
(1) hydrobromic acid = H^+ Br^- = HBr (2) acetic acid = H^+ $C_2H_3O_2^-$ = $HC_2H_3O_2$ (3) nitrous acid = H^+ NO_2^- = HNO_2

Covalent Compounds
- Form between non–metal atoms.
- Electrons are shared between atoms so that all atoms have a full valence shell.

I. Naming Covalent Compounds
(1) Name the first element and use a prefix to if there is more than one atom (if there is only one atom of the first element, it is not necessary to use the prefix “mono”). Do not change the ending of the first element.
(2) Name second element and use a prefix to indicate the number of atoms (if there is only one atom of the second element, the prefix "mono" must be used). Change the ending to “ide”.

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

ex. NI_3
nitrogen triiodide

ex. P_2O_5
diphosphorus pentoxide

ex. CO
carbon monoxide

ex. B_5H_9
pentaboron nonahydride

Dropping vowels between the ending of a prefix and the first letter of an element
The vowel at the end of a prefixes ending in “a” or “o” is dropped for compounds where the second element begins with a an “a” or “o” (ex. monooxide is written monoxide or pentaoxide is written pentoxide).
The vowel is not dropped if the prefix ends in an “i” or the second element begins with an “i” (ex. dioxide, triiodide, tetraiodide).

II. Writing formulas for Covalent Compounds
(1) Write the symbol for the first element.
(2) Write the symbol for the second element.
(3) Use the prefixes to determine the number of atoms of each element and write the formula indicating the number of atoms of each element with subscripts.
NOTE: in covalent compounds, the charge of an atom is NOT used since electrons are shared not transferred. Do NOT show charges when writing formulas for covalent compounds.

ex. sulphur hexafluoride
= SF_6

ex. tricarbon octahydride
= C_3H_8

ex. tetraphosphorus decasulphide
= P_4S_{10}

ex. iodine heptafluoride
= IF_7

Diatomic Elements
Diatomic elements: these elements have molecules containing two atoms. The molecule for these diatomic elements is named with just the name of the element (i.e. it is not correct to use to prefix "di" when naming these molecules) and when writing the formula for the molecule, two atoms of the element are always present.

Element	Formula
hydrogen	H_2
oxygen	O_2
fluorine	F_2
bromine	Br_2
iodine	I_2
nitrogen	N_2
chlorine	Cl_2

Just remember HOFBrINCl

(6) The Mole

The mole is concerned with counting and measuring matter. Consider a single molecule of a compound. It is important to be able to determine the number of atoms of each element in the compound. The number of atoms of each element can be found from subscripts in the chemical formula. If there are parenthesis in the formula, the number of atoms of an element is calculated by multiplying the number outside the parenthesis by the subscript for that element.

ex. P_4O_{10} 4 phosphorus, 10 oxygen

ex. $Al(NO_3)_3$ 1 aluminum, 3 nitrogen (3x1), 9 oxygen (3x3)

In a laboratory setting, chemists use very large number of atoms or molecules, so they count in **moles**.

1 mole = $6.02x10^{23}$ atoms or molecules

This number is also called Avogadro's number after the chemist Amedeo Avogadro (1776–1856). The value of Avogadro's number was calculated by determining the number of carbon atoms required for one mole of **carbon–12** atoms have a mass of exactly 12.00 g. The symbol for mole is **mol**.

Molar Mass

Molar mass: the mass in grams of one mole of an element or compound. Molar mass has units of g/mol.
Molar mass for elements can be found below the element on the periodic table (molar mass is equivalent to atomic mass)

ex. sulphur 32.06 g/mol

ex. silver 107.87 g/mol

The molar mass of a compound can be determined by adding the molar mass of all the elements in the compound, taking into account the number of atoms of each element. Do not round calculated values for molar mass.

ex. What is the molar mass of Na_2O?

= 2(22.99) + 16.00 = 61.98 g/mol

ex. What is the molar mass of $(NH_4)_2SO_4$?

= 2(14.01) + 8(1.008) + 32.06 + 4(16.00) = 132.144 g/mol

Mole Conversions

It is important for chemists to be able to determine how much of a given substance is present in terms of …
(1) moles (2) mass (3) number of particles (atoms or molecules)
and to be able to convert between these amounts.

To convert between mass and moles, use molar mass
ex. What is the mass of 1.5 moles of chromium?

Chromium has a molar mass of 52.00 g/mol, this can be written as the conversion fraction $\frac{52.00\,g}{mol}$

To find mass, multiply the given amount of moles by the molar mass conversion fraction in order to cancel the units of mol (the conversion fraction is written with g in the numerator and mol in the denominator).

$$1.5\ mol \times \frac{52.00\,g}{mol} = 78\,g$$

ex. how many moles are in 55.0 g of carbon dioxide?

Carbon dioxide has a formula of CO_2, therefore the molar mass must be calculated.
12.01 + 2(16.00) = 44.01 g/mol

To find moles, multiply the given amount of grams by the molar mass conversion fraction in order to cancel the units of g (the conversion fraction is written with mol in the numerator and g in the denominator).

$$55.0\ g \times \frac{mol}{44.01\,g} = 1.25\,mol$$

To convert between moles and particles, use Avogadro's number (1 mol = $6.02x10^{23}$ atoms or molecules)

Avogadro's number can be written as two different conversion fractions:

$$\frac{6.02\times10^{23}\ atoms\ or\ molecules}{mol} \quad or \quad \frac{mol}{6.02\times10^{23}\ atoms\ or\ molecules}$$

Note: atoms are used to express the number of particles for substances with only one atom of one element (ex. Cr, V, Fe, etc.)
molecules are used the express the number of particles for substances with more than one atom, and/or more than one element (ex. H_2O, F_2, NaCl, etc.)

ex. how many atoms are in 0.500 mol of vanadium?

To find atoms, multiply the given amount of moles by the Avogadro's number conversion fraction in order to cancel the units of mol (the conversion fraction is written with atoms in the numerator and mol in the denominator).

$$0.500\,mol \times \frac{6.02\times10^{23}\ atoms}{mol} = 3.01\times10^{23}\ atoms$$

ex. how many moles are in 1.5x10^{23} molecules of water?

To find moles, multiply the given amount of molecules by the Avogadro's number conversion fraction in order to cancel the units of molecules (the conversion fraction is written with mol in the numerator molecules in the denominator).

$$1.5\times10^{23}\ molecules \times \frac{mol}{6.02\times10^{23}\ molecules} = 0.25\,mol$$

To convert between mass and atoms/molecules, the moles must first be calculated.

ex. What is the mass of 8.25x10^{22} molecules of NaF?
Determine the molar mass: 22.99 + 19.00 = 41.99 g/mol

Calculate moles: $8.25\times10^{22}\ molecules \times \frac{mol}{6.02\times10^{23}\ molecules} = 0.137\,mol$

Calculate mass: $0.137\ mol \times \frac{41.99\,g}{mol} = 5.75\,g$

Note: The work can be simplified by combining the two conversion fractions into one step.

$$8.25\times10^{22}\ molecules \times \frac{mol}{6.02\times10^{23}\ molecules} \times \frac{41.99\,g}{mol} = 5.75\,g$$

ex. (a) How many molecules are contained in 0.850 kg of K_3PO_4? (b) How many atoms are there of EACH element?
(a) Determine the molar mass: 3(39.10) + 30.97 + 4(16.00) = 212.27

Convert to grams: $0.850\,kg \times \frac{1000\,g}{1kg} = 850\ g$

$$850\ g \times \frac{mol}{212.27\,g} \times \frac{6.02\times10^{23}\ molecules}{mol} = 2.41\times10^{24}\ molecules\ K_3PO_4$$

(b) To find the number of atoms of each element, use the calculated value of molecules and the subscript of each element in the chemical formula.
ex. 1 molecule K_3PO_4 = 3 atoms K = 1 atom P = 4 atoms O
The molecules can be multiplied by an appropriate conversion fraction (with atoms for the specific element in the numerator and molecules in the denominator).

$$2.41\times10^{24}\ molecules\ K_3PO_4 \times \frac{3\,atoms\ K}{1\,molecule\ K_3PO_4} = 7.23\times10^{24}\ atoms\ K$$

$$2.41\times10^{24}\ molecules\ K_3PO_4 \times \frac{1\,atom\ P}{1\,molecule\ K_3PO_4} = 2.41\times10^{24}\ atoms\ P$$

$$2.41\times10^{24}\ molecules\ K_3PO_4 \times \frac{4\,atoms\ O}{1\,molecule\ K_3PO_4} = 9.64\times10^{24}\ atoms\ O$$

Percent Composition

Percent composition: the total molar mass of one type of element in a compound divided by the total molar mass of the compound. Percent composition makes it easier to compare the elemental make up of different compounds.

Calculation of Percent Composition

Determine the molar mass of each individual element in the compound. Add to find the total molar mass of the entire compound. Divide the molar mass of each individual element in the compound by the total molar mass of the compound. Multiply the decimal by 100 in order to convert to a percent.

ex. What is the percent composition of each element in $SrCl_2$? Give answers to one decimal place.

$$\begin{array}{lll} Sr & 1\times 87.62 = 87.62 & 87.62/158.52\,x100\% = 55.3\%\,Sr \\ Cl & 2\times 35.45 = 70.90 & 70.90/158.52\,x100\% = 44.7\%\,Cl \\ & \overline{\quad 158.52} & \end{array}$$

ex. what is the percent composition of each element in CaC_2O_4? Give answers to one decimal place.

$$\begin{array}{lll} Ca & 1\times 40.08 = 40.08 & 40.08/128.10\,x100\% = 31.3\%\,Ca \\ C & 2\times 12.01 = 24.02 & 24.02/128.10\,x100\% = 18.8\%\,C \\ O & 4\times 16.00 = 64.00 & 64.00/128.10\,x100\% = 50.0\%\,O \\ & \overline{\quad 128.10} & \end{array}$$

ex. Percent composition for Hydrates

(a) What is the percent composition of each element in $Na_2SO_4 \cdot 10H_2O$ (sodium sulphate decahydrate)? Give answers to one decimal place. Note: the coefficient in front of water multiplies the subscript for both hydrogen and oxygen to determine the number of atoms of each of these elements.

$$\begin{array}{lll} Na & 2\times 22.99 = 45.98 & 45.98/322.2\,x100\% = 14.3\%\,Na \\ S & 1\times 32.06 = 32.06 & 32.06/322.2\,x100\% = 10.0\%\,S \\ O & 14\times 16.00 = 224 & 224/322.2\,x100\% = 69.5\%\,O \\ H & 20\times 1.008 = 20.16 & 20.16/322.2\,x100\% = 6.3\%\,H \\ & \overline{\quad 322.2} & \end{array}$$

(b) What is the percent composition of water in the compound?

To find the percent composition of water, find the molar mass for the given number of water molecules (i.e. 10) in the hydrate and divide by the total molar mass of the compound. Multiply the decimal by 100 in order to convert to a percent.

$10\ H_2O = 10(18.016) = 180.16$

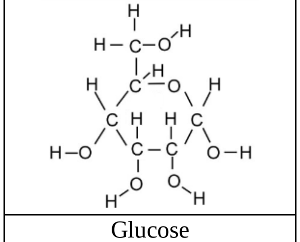
$$180.16/322.2\,x100\% = 55.9\%\,H_2O$$

Empirical and Molecular Formulas

The formula of a compound can be expressed in two ways: the empirical formula and the molecular formula.

Empirical formula: the simplest, reduced formula for a compound.
Molecular formula: the formula for the compound as it is actually found.

ex. Glucose has the molecular formula of $C_6H_{12}O_6$ (since this is the number of atoms if each element found in its structure) but can also be reduced (by dividing by six) to its empirical formula CH_2O.



Glucose

Note: The molecular formula is always a multiple of the empirical formula. As a result, the molar mass of the molecular formula (the molecular weight) is always a multiple of the molar mass of the empirical formula (the empirical weight).

ex.

	Empirical	**Molecular**
Formula	CH_2O x6	= $C_6H_{12}O_6$
Molar Mass	30.026 g/mol x6	= 180.156 g/mol

Determination of Empirical Formula
The empirical formula is determined from the mass or percent composition for a compound.

ex. Determine the empirical formula if 3.50 g of a compound is found to contain 2.21 g of manganese and 1.29 g of oxygen.
(1) Create a table with a row for each element and columns for mass, moles, and ratio.
(2) Convert the mass of each element to moles.
(3) Find the ratio of moles of each element by dividing the moles of each element by the smallest number of moles.
(4) Write the empirical formula. The ratio of moles indicate the number of atoms of each element (i.e. the subscripts) in the compound.

Element	Mass	Moles	Ratio
Mn	2.21 g	$2.21\,g\ x\ \frac{mol}{54.94\,g} = 0.0402\,mol$	$\frac{0.0402\,mol}{0.0402\,mol} = 1$
O	1.29 g	$1.29\,g\ x\ \frac{mol}{16.00\,g} = 0.0806\,mol$	$\frac{0.0806\,mol}{0.0402\,mol} \approx 2$

empirical formula = MnO_2

ex. Determine the empirical formula for a compound that contains 65.2% scandium and 34.8% oxygen.
When given percent composition set the total mass of the compound to 100 g, such that the percentages of each element are equivalent to mass. i.e. 65.2% of 100 g is 65.2 g and 34.8% of 100 g is 34.8 g. Calculate the moles of each element and the ratio of moles as before

Element	Mass	Moles	Ratio
Sc	65.2 g	$65.2\,g\ x\ \frac{mol}{44.96\,g} = 1.45\,mol$	$\frac{1.45\,mol}{1.45\,mol} = 1 \quad 1\ x\ 2 = 2$
O	34.8 g	$34.8\,g\ x\ \frac{mol}{16.00\,g} = 2.18\,mol$	$\frac{2.18\,mol}{1.45\,mol} \approx 1.5 \quad 1.5\ x\ 2 = 3$

empirical formula = Sc_2O_3, scandium oxide

Chemical formulas must be given with whole numbers. If one of the ratios does not turn out to be a whole number (generally within about 0.05), they must all be multiplied by the same number so that they are all whole numbers.

Decimal Place	Multiply by
____.25	4
____.33	3
____.5	2
____.67	3
____.75	4

Determination of Molecular Formula
Molecular formulas can be determined from the empirical formula and the molecular weight (the molar mass of the molecular formula).

ex. Paradichlorobenzene is the chemical found in moth balls. This compound consists of 49.0% carbon, 2.80% hydrogen, and 48.2% chlorine. The molecular weight is 146.992 g/mol. Determine the molecular formula.

(1) Determine the empirical formula
(2) Determine the molar mass for the empirical formula (the empirical weight).
(3) Determine the value of the given molecular weight divided by the empirical weight. This value indicates what to multiply the empirical formula by to obtain the molecular formula. This value will always be a whole number.

Element	Mass	Moles	Ratio
C	49.0 g	$49.0\,g\ x\ \frac{mol}{12.01\,g} = 4.08\,mol$	$\frac{4.08\,mol}{1.36\,mol} \approx 3$
H	2.80 g	$2.80\,g\ x\ \frac{mol}{1.008\,g} = 2.78\,mol$	$\frac{2.78\,mol}{1.36\,mol} \approx 2$
Cl	48.2 g	$48.2\,g\ x\ \frac{mol}{35.45\,g} = 1.36\,mol$	$\frac{1.36\,mol}{1.36\,mol} = 1$

empirical formula = C_3H_2Cl

empirical weight = 3(12.01) + 2(1.008) + 35.45 = 73.496 g/mol

$$\frac{molecular\ weight}{empirical\ weight} = \frac{146.992}{73.496} = 2$$

molecular formula = C_3H_2Cl x 2 = $C_6H_4Cl_2$

Molecular Formulas and Combustion Reactions, Combustion Analysis
The molecular formula of an organic compound can also be determined by combustion analysis– the compound is burned in oxygen and the masses of the products (carbon dioxide and water) are used to determine the amount of each element in the original sample.
ex. A compound contains only carbon, hydrogen, and oxygen. A 0.7549 g sample of the compound burns in oxygen to produce 1.9060 g of carbon dioxide and 0.3370 g of water.
(a) Determine the mass of each element in the sample.
(b) Determine the empirical formula.
(c) The molecular weight of the compound is 244.236 g/mol. Determine the molecular formula.

The mass of carbon can be found from the mass of carbon dioxide:

$$1.9061\, g\, CO_2\, x \frac{1\, mol\, CO_2}{44.01\, g\, CO_2}\, x \frac{1\, mol\, C}{1\, mol\, CO_2}\, x \frac{12.01\, g\, C}{1\, mol\, C} = 0.52016\, g\, C$$

The mass of hydrogen can be found from the mass of water:

$$0.3370\, g\, H_2O\, x \frac{1\, mol\, H_2O}{18.016\, g\, H_2O}\, x \frac{2\, mol\, H}{1\, mol\, H_2O}\, x \frac{1.008\, g\, H}{1\, mol\, H} = 0.03771\, g\, H$$

mass oxygen = mass sample – mass carbon – mass hydrogen = 0.7549 g –0.52016 g – 0.03771 g = 0.1970 g O

Element	**Mass**	**Moles**	**Ratio**
C	0.52016	$0.52016\, g\, x \frac{mol}{12.01\, g} = 0.043311\, mol$	$\frac{0.043311\, mol}{0.01231\, mol} \approx 3.5 \quad 3.5\, x\, 2 = 7$
H	0.03771	$0.03771\, g\, x \frac{mol}{1.008\, g} = 0.03741\, mol$	$\frac{0.03741\, mol}{0.01231\, mol} \approx 3 \quad 3\, x\, 2 = 6$
O	0.1970	$0.1970\, g\, x \frac{mol}{16.00\, g} = 0.01231\, mol$	$\frac{0.01231\, mol}{0.01231\, mol} = 1 \quad 1\, x\, 2 = 2$

empirical formula = $C_7H_6O_2$

empirical weight = 7(12.01) + 6(1.008) + 2(16.00) = 122.118 g/mol

$$\frac{molecular\ weight}{empirical\ weight} = \frac{244.236}{122.118} = 2$$

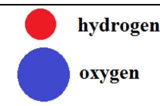
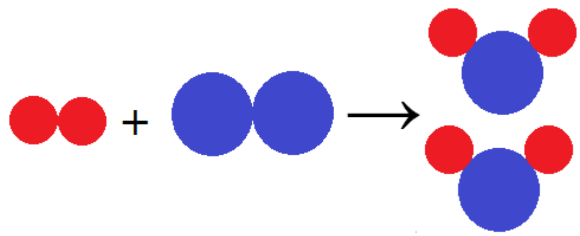
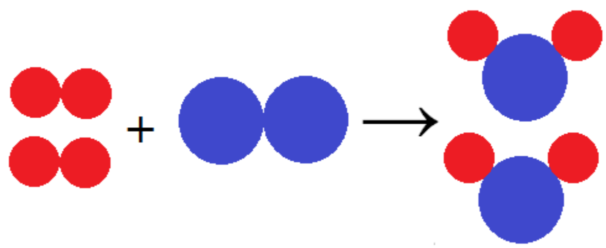
molecular formula = $C_7H_6O_2$ x 2 = $C_{14}H_{12}O_4$

(7) Chemical Reactions

Law of Conservation of Mass

Matter cannot be created or destroyed, therefore, the number of atoms of each element on the reactants side of a chemical equation must be equal to the number of atoms of each element on the products side of a chemical equation. Furthermore, the mass of the reactants must be equal to the mass of the products. Chemical reactions must always be given as balanced equations.

Example:

Chemical Equation	**Representation** 
hydrogen + oxygen → water H_2 + O_2 → H_2O Initially, there are 2 hydrogen atoms on 2 oxygen atoms on the left side of the reaction and 2 hydrogen atoms and 1 oxygen atom on the right side of the reaction, so the equation is not balanced because the number of oxygen atoms is not equal.	
H_2 + O_2 → **2**H_2O A 2 is added in front of H_2O on the right in order to balance the oxygen atoms, such that there are now there are 2 hydrogen atoms on 2 oxygen atoms on the left side of the reaction and 4 hydrogen atoms and 2 oxygen atoms on the right side of the reaction, so the equation is not balanced because the number of hydrogen atoms is not equal.	
2H_2 + O_2 → $2H_2O$ A 2 is added in front of H_2 on the left in order to balance the hydrogen atoms, such that there are now there are 4 hydrogen atoms on 2 oxygen atoms on the left side of the reaction and 4 hydrogen atoms and 2 oxygen atoms on the right side of the reaction, so **the equation is balanced** because the number of atoms of both elements are equal.	

Rules for Balancing

No new reactants or products can be added (i.e. + O cannot be added on the products side to balance the oxygen atoms).

The subscripts cannot be changed (because this would change what the substance actually is, i.e. H_2O is water and H_2O_2 is hydrogen peroxide).

Only the coefficients (numbers in front of the substance) can be changed.

The coefficients represent that ratio in which substances react or are produced (i.e. **two** molecules of hydrogen react with **one** molecule of oxygen to produce **two** molecules of water).

The coefficients are often listed for the balanced chemical reaction (i.e. 2, 1, 2).

Examples:

(1) zinc oxide → zinc + oxygen

$2ZnO \rightarrow 2Zn + O_2$

(2) aluminum + iron (III) oxide → aluminum oxide + iron

$2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe$

(3) calcium nitrate + sodium phosphate → calcium phosphate + sodium nitrate
(for balancing, treat polyatomic ions as a unit)

$3Ca(NO_3)_2 + 2Na_3PO_4 \rightarrow Ca_3(PO_4)_2 + 6NaNO_3$

Reaction Types

(1) Synthesis
A + B → AB
example: $2H_2 + O_2 \rightarrow 2H_2O$

(2) Decomposition
AB → A + B
example: $2NaCl \rightarrow 2Na + Cl_2$

(3) Single Replacement
AB + C → AC + B *OR* AB + C → CB + A
* metals replace metals and non–metals replace non–metals!

example: $2MgO + 2Cl_2 \rightarrow 2MgCl_2 + O_2$
(here, the non–metal chlorine replaces the non–metal oxygen)

example: $CaS + 2Ag \rightarrow Ag_2S + Ca$
(here, the metal silver replaces the metal calcium)

(4) Double Replacement
AB + CD → AD + CB
(both the metals and non–metals switch places)
example: $CaCl_2 + Na_2CO_3 \rightarrow CaCO_3 + 2NaCl$

(5) Neutralization
acid + base → salt + water
acid: begins with an H
base: ends with an OH
salt: any ionic compound
water: H_2O– can be written at HOH
example: $HCl + NaOH \rightarrow NaCl + H_2O$

(6) Combustion
hydrocarbon + oxygen → carbon dioxide + water
hydrocarbon: contains hydrogen and carbon
example: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
Balance using "CHO-CHO" (balance carbon, then hydrogen, then oxygen last)

Predicting Products

To predict the products of a reaction, first determine the type of reaction.
(1) **Synthesis:** there will be two reactants which are both elements. There will be only one product. To determine the formula of the product, combine the elements and balance their charges.
(2) **Decomposition:** There will be only one reactant which is a compound. There will be two products which are both elements. Remember HOFBrINCl, if applicable, when writing the formulas for products.
(3) **Single Replacement:** There will be two reactants– one is a compound and the other is an element by itself. There will be two products– again, one will be a compound and the other will be an element by itself. Classify each element in the reactants as a metal or non–metal to determine which elements will switch. If there are two metals, they will switch places (so there will be a metal by itself in the products), if there are two non–metals, they will switch places (so there will be a non–metal by itself in the products). Determine the formula of the product that is the compound by balancing the charges of the elements that combine together. In a compound, the metal is always written first followed by the non–metal.
(4) **Double Replacement:** There will be two reactants which are both compounds. There will be two products which are both compounds. Here, the metals switch places with each other and the nonmetals switch places with each other. The easy rule is "outsides together, insides together". Be sure the balance the charges for each compound produced and remember the metal is always written first followed by the non–metal. If there is a multivalent metal in the reaction, the charge can be determined from the charge on the reactants side.
(5) **Neutralization:** The reactants will be an acid (starts with H) and a base (ends with OH). One of the products will always be water (HOH), the other product is a salt. The formula for the salt can be written by taking the beginning of the base and the ending of the acid and balancing the charges.
(6) **Combustion:** The reactants will be a hydrocarbon and oxygen and the products are always carbon dioxide (CO_2) and water (H_2O).
Note: if there is more than one product, the order the products are written is not important.

Practice Examples: classify each reaction, predict the products, and balance the chemical equation

______________ (1) $Mg + O_2 \rightarrow$

______________ (2) $C_9H_{20} + O_2 \rightarrow$

______________ (3) $HCl + Sr(OH)_2 \rightarrow$

______________ (4) $NCl_3 \rightarrow$

______________ (5) $KCl + Pb(NO_3)_2 \rightarrow$

______________ (6) $ZnCl_2 + Na \rightarrow$

______________ (7) $CaBr_2 + O_2 \rightarrow$

Answers:
(1) Synthesis: $2Mg + O_2 \rightarrow 2MgO$
(2) Combustion: $C_9H_{20} + 14O_2 \rightarrow 9CO_2 + 10H_2O$
(3) Neutralization: $2HCl + Sr(OH)_2 \rightarrow SrCl_2 + 2H_2O$
(4) Decomposition: $2NCl_3 \rightarrow N_2 + 2Cl_2$
(5) Double Replacement: $2KCl + Pb(NO_3)_2 \rightarrow 2KNO_3 + PbCl_2$
(6) Single Replacement: $ZnCl_2 + 2Na \rightarrow Zn + 2NaCl$
(7) Single Replacement: $2CaBr_2 + O_2 \rightarrow 2CaO + 2Br_2$

Summary of Basic AP Reactions

Synthesis Reactions
Metals combine with nonmetals to form a salt (any ionic compound)
ex. $6Li + 3N_2 \rightarrow 2Li_3N$

***Metallic oxides react with water to form bases (metallic hydroxides)
ex. $Na_2O + H_2O \rightarrow 2NaOH$ ex. $MgO + H_2O \rightarrow Mg(OH)_2$

***Nonmetallic oxides react with water to form acids
ex. $CO_2 + H_2O \rightarrow H_2CO_3$ ex. $SO_3 + H_2O \rightarrow H_2SO_4$

***Metallic oxides and nonmetallic oxides react to form salts
ex. $Na_2O + CO_2 \rightarrow Na_2CO_3$ ex. $CaO + SO_3 \rightarrow CaSO_4$

***Alkali metals react with water to form a base and hydrogen gas
ex. $2Na + 2H_2O \rightarrow 2NaOH + H_2$

Decomposition Reactions
Salts decompose into metals and nonmetals
ex. $2NaCl \rightarrow 2Na + Cl_2$

***Metallic carbonates decompose into metallic oxides and carbon dioxide
ex. $CaCO_3 \rightarrow CaO + CO_2$

***Metallic chlorates decompose into metallic chlorides and oxygen
ex. $2KClO_3 \rightarrow 2KCl + 3O_2$

***Ammonium carbonate decomposes into ammonia, water, and carbon dioxide
$(NH_4)_2CO_3 \rightarrow 2NH_3 + H_2O + CO_2$

***Ammonium hydroxide decomposes into ammonia and water
$NH_4OH \rightarrow NH_3 + H_2O$

***Sulphurous acid decomposes into sulphur dioxide and water
$H_2SO_3 \rightarrow SO_2 + H_2O$

***Carbonic acid decomposes into carbon dioxide and water
$H_2CO_3 \rightarrow CO_2 + H_2O$

***Hydrogen peroxide decomposes into water and oxygen
$2H_2O_2 \rightarrow 2H_2O + O_2$

***memorize and be able to apply these reaction patterns

Practice Examples

(1) $CaO + H_2O \rightarrow$

(2) $SO_2 + H_2O \rightarrow$

(3) $MgO + CO_2 \rightarrow$

(4) $Na_2CO_3 \rightarrow$

(5) $Ca(ClO_3)_2 \rightarrow$

(6) $K + H_2O \rightarrow$

Answers:
(1) $CaO + H_2O \rightarrow Ca(OH)_2$
(2) $SO_2 + H_2O \rightarrow H_2SO_3$
(3) $MgO + CO_2 \rightarrow MgCO_3$
(4) $Na_2CO_3 \rightarrow Na_2O + CO_2$
(5) $Ca(ClO_3)_2 \rightarrow CaCl_2 + 3O_2$
(6) $2K + 2H_2O \rightarrow 2KOH + H_2$

(8) Stoichiometry

Mole Ratios
The moles of substances in a chemical reaction are always in proportion as determined from the balanced chemical equation.

ex. Given the balanced reaction: $O_2 + 4HBr \rightarrow 2H_2O + 2Br_2$

From the coefficients of the equation, it can be seen that the chemicals must be combined in the following proportions: for every one mole of O_2 reacting there must be four moles of HBr reacting and there will be two moles of H_2O produced and two moles of Br_2 produced.

$$1 \text{ mol } O_2 = 4 \text{ mol HBr} = 2 \text{ mol } H_2O = 2 \text{ mol } Br_2$$

Given the moles of one substance in a reaction, the moles of any other substance can be determined by multiplying by a mole ratio. A mole ratio is a conversion fraction that shows the ratio of coefficients of two substances from the balanced chemical equation. The conversion fraction is written such that the coefficient of the given substance is in the denominator and the coefficient of the substance to be determined is in the numerator (in order for the units to cancel appropriately).

ex. if 2.00 moles of HBr react, determine the moles of O_2 reacting.

$$2.00\, mol\ HBr\ x \underbrace{\frac{1\, mol\ O_2}{4\, mol\ HBr}}_{\text{MOLE RATIO}} = 0.500\, mol\ O_2$$

ex. if 0.080 moles of H_2O are produced, determine the moles of HBr reacting

$$0.080\, mol\ H_2O\ x \frac{2\, mol\ HBr}{2\, mol\ H_2O} = 0.16\, mol\ HBr$$

ex. if 15 mol of Br_2 are produced, determine the moles of O_2 reacting

$$15\, mol\ Br_2\ x \frac{1\, mol\ O_2}{2\, mol\ Br_2} = 7.5\, mol\ O_2$$

Stoichiometry Calculations involving Mass
ex. Sodium fluoride is reacted with calcium nitrate. If 8.40 g of sodium fluoride are present, what mass of calcium nitrate would be required in the reaction? What would be the mass of each of the products?
(1) Determine a balanced chemical equation: Write the chemical formulas for each substance, being sure to balance the charges if needed. Add coefficients so that the number of atoms of each element is the same of both side of the reaction.
(2) Draw a table with three rows labeled moles (mol), mass (g), and molar mass (g/mol) with a column below each chemical. (see below)
(3) Determine the given information and enter it below the proper chemical in the mass row.
(4) Calculate the molar mass of each chemical (remember, coefficients are not used in calculating molar mass).
(5) Use the given mass to calculate the number of moles present.
(6) Use the moles that were calculated to determine the moles of other chemicals in the reaction, using the correct mole ratio.
(7) Use the moles to determine the mass each chemical.
(Keep numbers in your calculator for intermediate calculations)

Reaction:	2NaF +	$Ca(NO_3)_2$ →	$2NaNO_3$ +	CaF_2
moles (mol)	0.200	0.100	0.200	0.100
mass (g)	**8.40** (given)	16.4	17.0	7.81
molar mass (g/mol)	22.99 + 19.00 = 41.98	40.08 + 2(14.01) + 6(16.00) = 164.10	22.99 + 14.01 + 3(16.00) = 85.00	40.08 + 2(19.00) = 78.08

$$8.40 \text{ g x } \frac{\text{mol}}{41.98 \text{ g}} = 0.200 \text{ mol NaF}$$

$$0.200 \text{ mol NaF x } \frac{1 \text{ mol } Ca(NO_3)_2}{2 \text{ mol NaF}} = 0.100 \text{ mol } Ca(NO_3)_2 \qquad 0.100 \text{ mol x} \frac{164.10 \text{ g}}{\text{mol}} = 16.4 \text{ g } Ca(NO_3)_2$$

$$0.200 \text{ mol NaF x } \frac{2 \text{ mol } NaNO_3}{2 \text{ mol NaF}} = 0.200 \text{ mol } NaNO_3 \qquad 0.200 \text{ mol x} \frac{85.00 \text{ g}}{\text{mol}} = 17.0 \text{ g } NaNO_3$$

$$0.200 \text{ mol NaF x } \frac{1 \text{ mol } CaF_2}{2 \text{ mol NaF}} = 0.100 \text{ mol } CaF_2 \qquad 0.100 \text{ mol x} \frac{78.08 \text{ g}}{\text{mol}} = 7.81 \text{ g } CaF_2$$

Note: Check answers using the law of conservation of mass.
mass reactants = mass products
8.40 g + 16.4 g = 17.0 g + 7.81 g
24.8 g = 24.8 g √ correct

Percent Yield

In laboratory stoichiometry experiments, the amount of product predicted by stoichiometry is often not obtained. Side reactions, contamination, incomplete reaction, or loss of product can cause the amount of product actually obtained (the "actual mass") to be lower than the amount of product predicted by stoichiometry (the "theoretical mass"). Percent Yield is calculated by dividing the actual mass a given product by the theoretical mass of that product.

$$\text{Percent yield} = \frac{\text{actual mass}}{\text{theoretical mas}} \text{ x } 100\%$$

Percent yield will always be equal to or less than 100%. Percent yield can vary for a reaction depending on the reaction conditions (i.e. different experimental conditions can increase or decrease the percent yield). The actual yield for other products in a reaction can be calculated using the percent yield and the theoretical mass.

ex. Aluminum sulphide can be decomposed to form aluminum and sulphur. If 24.0 g of aluminum sulphide reacts, calculate the theoretical yield of each of the products. If the reaction actually produces 7.40 g of aluminum, determine the percent yield for the reaction. What mass of sulphur would actually be obtained?

Reaction: $Al_2S_3 \rightarrow 2Al + 3S$

	Al_2S_3	2Al	3S
moles (mol)	0.160	0.320	0.480
mass (g)	**24.0** (given)	8.63	15.4
molar mass (g/mol)	2(26.98) + 3(32.06) = 150.14	26.98	32.06

$$24.0 \text{ g x } \frac{\text{mol}}{150.14 \text{ g}} = 0.160 \text{ mol } Al_2S_3$$

$$0.160 \text{ mol } Al_2S_3 \text{ x } \frac{2 \text{ mol Al}}{1 \text{ mol } Al_2S_3} = 0.320 \text{ mol Al} \qquad 0.320 \text{ mol Al x } \frac{26.98 \text{ g}}{\text{mol}} = 8.63 \text{ g Al (theoretical mass)}$$

$$0.160 \text{ mol } Al_2S_3 \text{ x } \frac{3 \text{ mol S}}{1 \text{ mol } Al_2S_3} = 0.480 \text{ mol S} \qquad 0.480 \text{ mol S x } \frac{32.06 \text{ g}}{\text{mol}} = 15.4 \text{ g S (theoretical mass)}$$

$$\text{Percent yield} = \underbrace{\frac{\text{actual mass}}{\text{theoretical mas}}}_{\text{(for aluminum)}} \text{ x } 100\% = \frac{7.40 \text{ g}}{8.63 \text{ g}} \text{ x } 100\% = 85.7\%$$

$$\text{actual mass (for sulphur)} = (0.857)(15.4 \text{ g}) = 13.2 \text{ g S (actual mass)}$$

Limiting and Excess Reactions

It is common when preparing a chemical reaction to add more of one of the reactants than is actually necessary (more than would be required by stoichiometry). Adding extra of one reactant can help to ensure that the other reactant is entirely used up. In this situation. the reactant with a greater amount of moles would be in **excess**, so it is called the **excess reactant** and the reactant with a smaller amount of moles would be **limiting**, so it is called the **limiting reactant**.

If the mass of both reactants are given, begin by calculating the moles of each reactant. Determine which reactant is limiting and which is in excess (by dividing the moles of each reactant by the coefficient and finding which is lower). Use the moles of the limiting reactant to calculate the moles and mass of product formed. To determine the mass of the excess reactant used in the reaction, use the moles of the limiting reactant to calculate the moles of the excess reactant used and then the mass of the excess reactant used. To calculate the mass of the excess reactant remaining, subtract the mass of the excess reactant used from the mass of the excess reactant available/given.

ex. Zinc reacts with phosphorus to produce zinc phosphide.
51.0 g of zinc are reacted with 19.2 g of phosphorus.

(a) Which reactant is limiting and which is excess?
(b) What is the theoretical mass of the product?
(c) What mass of the excess reactant is used in the reaction and what mass of the excess reactant remains after the reaction?
(d) Calculate the actual mass of product obtained if the reaction has an 80.0% yield.

Reaction:	3Zn +	2P →	Zn_3P_2
moles (mol)	0.780	0.620	0.260
mass (g)	**51.0** (given)	**19.2** (given)	67.1
molar mass (g/mol)	65.39	30.97	3(65.39) + 2(30.97) 258.11

$51.0 \text{ g x } \frac{\text{mol}}{65.38 \text{ g}} = 0.780 \text{ mol Zn}$ $0.780/3 = 0.260$ Zn is LIMITING

$19.2 \text{ g x } \frac{\text{mol}}{30.97 \text{ g}} = 0.620 \text{ mol P}$ $0.620/2 = 0.310$ P is EXCESS

$0.780 \text{ mol Zn x } \frac{1 \text{ mol } Zn_3P_2}{3 \text{ mol Zn}} = 0.260 \text{ mol } Zn_3P_2$ $0.260 \text{ mol } Zn_3P_2 \text{ x } \frac{258.11 \text{ g}}{\text{mol}} = 67.1 \text{ g } Zn_3P_2$ (theoretical mass)

$0.780 \text{ mol Zn x } \frac{2 \text{ mol P}}{3 \text{ mol Zn}} = 0.520 \text{ mol P (used)}$ $0.520 \text{ mol P x } \frac{30.97 \text{ g}}{\text{mol}} = 16.1 \text{ g P used}$

$\text{mass remaining} = 19.2 \text{ g} - 16.1 \text{ g} = 3.1 \text{ g P remaining}$

$\text{actual mass} = (0.800)(67.1 \text{ g}) = 53.7 \text{ g } Zn_3P_2$

(9) Gases

An important property of a gas is pressure.

Pressure: the amount of force exerted on a surface due to collisions by gas particles. There are several units that are used to express the amount of pressure including atmospheres (atm), millimeters of mercury (mm Hg), and kilopascals (kPa).

1.00 atm = 760 mm Hg (equivalent to torr) = 101.3 kPa

ex. 0.850 atm = __________ mm Hg (torr) = __________ kPa = __________ Pa

$$0.850\, atm\, x \frac{760\, mmHg}{1 atm} = 646\, mmHg\ or\ 646\, torr$$

$$0.850\, atm\, x \frac{101.3 kPa}{1 atm} = 86.1 kPa \quad 86.1 kPa\, x \frac{1000\, Pa}{1 kPa} = 8.61 x 10^4\ Pa$$

Gas Laws

(1) Dalton's Law

The total pressure of a mixture of gases is equal to the sum of the partial pressures of all the gases in the mixture

$$P_{Total} = P_1 + P_2 + P_3 \ldots$$

ex. A mixture contains nitrogen with a partial pressure of 5.5 atm and oxygen with a partial pressure of 6.5 atm, what is the total pressure of the mixture?

$$P_{Total} = P_{Nitrogen} + P_{oxygen} = 5.5\, atm + 6.5\, atm = 12.0\, atm$$

(2) Boyle's Law
At a constant temperature, Boyle's Law gives the relationship between changing pressure/volume.

$P_1V_1 = P_2V_2$

P_1 = initial pressure P_2 = final pressure
V_1 = initial volume V_2 = final volume

Pressure can be any consistent unit.
Volume can be any consistent unit.

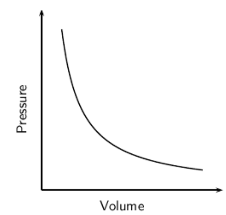
ex. A sample of argon gas at 0.950 atm of pressure occupies a volume of 1.50 L. What is the volume if the pressure is increased to 1.25 atm?

$P_1V_1 = P_2V_2$
$(0.950 \text{ atm})(1.50 \text{ L}) = (1.25 \text{ atm})V_2$
$V_2 = 1.14 \text{ L}$

ex. A sample of air under 150 kPa of pressure has a volume of 175 mL. What is the pressure if the volume is increased to 200 mL?

$P_1V_1 = P_2V_2$
$(150 \text{ kPa})(175 \text{ mL}) = P_2(200 \text{ mL})$
$P_2 = 131 \text{ kPa}$

When pressure is increased, volume is **decreased**.
When volume is increased, pressure is **decreased**.
Pressure and volume are INVERSELY proportional.



(3) Charles' Law
At a constant pressure, Charles' Law gives the relationship between changing volume/temperature.

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$

V_1 = initial volume V_2 = final volume
T_1 = initial temperature T_2 = final temperature

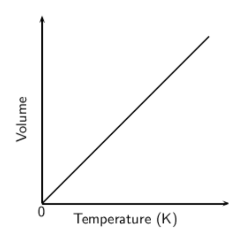
Volume can be any consistent unit.
TEMPERATURE MUST BE IN KELVIN

ex. A sample of oxygen gas occupies 0.750 L at 300 K. What volume with the gas occupy if the temperature is increased to 400 K?

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
$\frac{0.750 \text{ L}}{300 \text{ K}} = \frac{V_2}{400 \text{ K}}$
$V_2 = 1.00 \text{ L}$

ex. A sample of chlorine gas occupies 80.0 mL at 30 °C. At what temperature will the volume be 50.0 mL?
$T_1 = 30.0 + 273 = 303 \text{ K}$
$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
$\frac{80.0 \text{ mL}}{303 \text{ K}} = \frac{50.0 \text{ mL}}{T_2}$
$T_2 = 189 \text{ K}$

When temperature is increased, volume is **increased**.
When volume is decreased, temperature is **decreased**.
Pressure and volume are DIRECTLY proportional.



(4) Gay-Lussac's Law
At constant volume, Gay-Lussac's Law gives the relationship between changing pressure/temperature.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

P_1 = initial pressure $\quad$ P_2 = final pressure
T_1 = initial temperature $\quad$ T_2 = final temperature

Pressure can be any consistent unit.
TEMPERATURE MUST BE IN KELVIN

ex. A sample of nitrogen gas at 1.2 atm of pressure and 275 K is heated to 350 K. What is the resulting pressure?

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$\frac{1.2\ atm}{275\ K} = \frac{P_2}{350\ K}$$

$$P_2 = 1.5\ atm$$

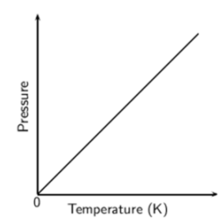
ex. A sample of xenon is under of pressure of 1.30×10^5 Pa at 50 °C. What is the temperature when the pressure is decreased to 8.90×10^4 Pa?
T_1 = 50 + 273 = 323 K

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$\frac{1.30x10^5 Pa}{323\ K} = \frac{8.90x10^4 Pa}{T_2}$$

$$T_2 = 221\ K$$

When temperature is increased, pressure is **increased**.
When pressure is decreased, temperature is **decreased**.
Pressure and temperature are DIRECTLY proportional.



(5) Combined Gas Law
Gives the relationship between changing pressure/volume/temperature for a gas.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

P_1 = initial pressure $\quad$ P_2 = final pressure
V_1 = initial volume $\quad$ V_2 = final volume
T_1 = initial temperature $\quad$ T_2 = final temperature

Pressure and Volume can be any consistent unit.
TEMPERATURE MUST BE IN KELVIN

ex. A balloon filled with helium has a volume of 4.0 L at 20 °C and 1.04 atm. What is the volume at 25 °C and 0.90 atm?

T_1 = 20 + 273 = 293 K $\quad$ T_2 = 25 + 273 = 298 K

$$\frac{\mathrm{P_1V_1}}{\mathrm{T_1}} = \frac{\mathrm{P_2V_2}}{\mathrm{T_2}}$$

$$\frac{(1.04\ \mathrm{atm})(4.0\ \mathrm{L})}{293\ \mathrm{K}} = \frac{(0.90\ \mathrm{atm})\mathrm{V_2}}{298\ \mathrm{K}}$$

$$\mathrm{V_2} = 4.7\ \mathrm{L}$$

(6) Avogadro's Law
At STP = Standard Temperature (0 °C or 273 K) and Pressure (1.00 atm or 101.3 kPa), 1 mole of ANY gas will occupy 22.4 L.
1 mole = 22.4 L

ex. What is the mass of 2.00 L of Neon at STP? How many atoms are present?

$$2.00\ L\ x\ \frac{mol}{22.4\ L}\ x\ \frac{20.18\ g}{mol} = 1.80\ g\ Ne$$

$$2.00\ L\ x\ \frac{mol}{22.4\ L}\ x\ \frac{6.02x10^{23} atoms}{mol} = 5.38x10^{22} atoms\ Ne$$

ex. How many molecules of nitrogen trifluoride are contained in 125 mL at STP?

$$125\ mL\ x\ \frac{1\ L}{1000\ mL}\ x\ \frac{mol}{22.4\ L}\ x\ \frac{6.02x10^{23} molecules}{mol} = 3.36x10^{21} molecules\ NF_3$$

(7) Ideal Gas Law
Gives the relationship between pressure/volume/temperature/moles of gas for any condition.

$PV = nRT$

P = pressure (atm or kPa)
V = volume (L)
n = moles (mol)
R = 0.0821 L atm/mol K or 8.314 L kPa/mol K (use units of given pressure to decide which value of R to use)
T = temperature (K)

ex. What is the volume of a sample containing 0.540 mol of nitrogen gas at 72 ºC and $1.15x10^5$ Pa?
$K = C + 273 = 72 + 273$
$= 345\ K\ (note: the\ significant\ figures\ for\ the\ value\ of\ temperature\ used\ in\ the\ ideal\ gas\ law\ would\ be\ three)$

$$1.15x10^5\ Pa\ x \frac{1 kPa}{1000\ Pa} = 115\ kPa$$

$$PV = nRT$$
$$(115\ kPa)V = (0.540\ mol)(8.314\ LkPa/mol\ K)(345\ K)$$
$$V = 13.5\ L$$

ex. What is the mass of a 54 L sample of helium at 280 K and 0.85 atm.

$$PV = nRT$$
$$(0.85\ atm)(54\ L) = n(0.0821\ L\ atm/mol\ K)(280\ K)$$
$$n = 2.0\ mol$$
$$2.0\ mol\ x \frac{4.003\ g}{mol} = 8.0\ g\ He$$

Molar Mass and Vapour Density
ex. An unknown monatomic gas has a vapour density of 0.195 g/L at 300 K and 1.20 atm. Calculate the molar mass of the gas and identify the gas.

(1) Assume that the volume of the gas is 1.00 L.

(2) Calculate the mass of the gas when V = 1.00 L.
$m = DV = (0.195\ g/L)(1.00\ L) = 0.195\ g$

(3) Calculate the moles present when V = 1.00 L

$$\mathrm{n} = \frac{\mathrm{PV}}{\mathrm{RT}} = \frac{(1.20\ \mathrm{atm})(1.00\ \mathrm{L})}{(0.0821\ L\ atm/mol\ K)(300\ \mathrm{K})} = 0.0487\ mol$$

(4) Calculate the molar mass and identify the gas.
$molar\ mass = \frac{0.195\ g}{0.0487\ mol} = 4.00\ g/mol$

The unknown gas is helium.

Gas Stoichiometry

ex. Nitrogen and hydrogen combine to form ammonia (NH_3).
What volume and mass of nitrogen and hydrogen (at STP) would produce 8.52 g of ammonia?

	N_2 +	$3H_2$ →	$2NH_3$
moles (mol)	0.250	0.750	0.500
mass (g)	7.01	1.51	**8.52** (given)
volume (L)	5.60	16.8	
molar mass (g/mol)	2(14.01) = 28.02	2(1.008) =2.016	14.01 + 3(1.008) =17.034

$$8.52 \text{ g x } \frac{\text{mol}}{17.034 \text{ g}} = 0.500 \text{ mol NH}_3$$

$$0.500 \text{ mol NH}_3 \text{ x } \frac{1 \text{ mol N}_2}{2 \text{ mol NH}_3} = 0.250 \text{ mol N}_2 \qquad 0.250 \text{ mol N}_2 \text{ x} \frac{28.02 \text{ g}}{\text{mol}} = 7.01 \text{ g N}_2 \qquad 0.250 \text{ mol N}_2 \text{ x } \frac{22.4 \text{ L}}{\text{mol}} = 5.60 \text{ L N}_2$$

$$0.500 \text{ mol NH}_3 \text{ x } \frac{3 \text{ mol H}_2}{2 \text{ mol NH}_3} = 0.750 \text{ mol H}_2 \qquad 0.750 \text{ mol H}_2 \text{ x } \frac{2.016 \text{ g}}{\text{mol}} = 1.51 \text{ g H}_2 \qquad 0.750 \text{ mol H}_2 \text{ x } \frac{22.4 \text{ L}}{\text{mol}} = 16.8 \text{ L H}_2$$

ex. Propane (C_3H_8) is combusted with oxygen to produce carbon dioxide and water vapour.
(a) If 14.0 L of propane and 105 L of oxygen and are present at 1.02 atm and 17 °C, which reactant is limiting and which is excess?
(b) What would be the volume and mass of each of the products? (Assume the pressure and temperature remain constant)
(c) If the reaction actually produces 64.0 g of carbon dioxide, what is the percent yield? How much water vapour is produced?

	C_3H_8 +	$5O_2$ →	$3CO_2$ +	$4H_2O$
moles (mol)	0.600	4.50	1.80	2.40
mass (g)			79.2	43.2
volume (L)	**14.0 (given)**	**105 (given)**	42.0	56.0
molar mass (g/mol)			12.01 + 2(16.00) = 44.01	2(1.008) + 16.00 = 18.016

$$\text{n} = \frac{\text{PV}}{\text{RT}} \qquad \text{T} = 17 + 273 = 290 \text{ K}$$

$$\text{n}_{\text{C}_3\text{H}_8} = \frac{(1.02 \text{ atm})(14.0 \text{ L})}{(0.0821 \text{ Latm/mol K})(290 \text{ K})} = 0.600 \text{ mol C}_3\text{H}_8 \qquad \text{C}_3\text{H}_8 \text{ is LIMITING}$$

$$\text{n}_{\text{O}_2} = \frac{(1.02 \text{ atm})(105 \text{ L})}{(0.0821 \text{ Latm/mol K})(290 \text{ K})} = 4.50 \text{ mol O}_2 \qquad 4.50/5 = 0.900 \qquad \text{O}_2 \text{ is EXCESS}$$

$$\mathbf{0.600 \text{ mol C}_3\text{H}_8 \text{ x } \frac{3 \text{ mol CO}_2}{1 \text{ mol C}_3\text{H}_8} = 1.80 \text{ mol CO}_2}$$

$$\mathbf{1.80 \text{ mol x } \frac{44.01 \text{ g}}{\text{mol}} = 79.2 \text{ g CO}_2 \qquad V = \frac{(1.80 \text{ mol})(0.0821 \text{ Latm/mol K})(290 \text{ K})}{1.02 \text{ atm}} = 42.0 \text{ L CO}_2}$$

$$\mathbf{0.600 \text{ mol C}_3\text{H}_8 \text{ x } \frac{4 \text{ mol H}_2\text{O}}{1 \text{ mol C}_3\text{H}_8} = 2.40 \text{ mol H}_2\text{O}}$$

$$\mathbf{2.40 \text{ mol x } \frac{18.016 \text{ g}}{\text{mol}} = 43.2 \text{ g H}_2\text{O} \qquad V = \frac{(2.40 \text{ mol})(0.0821 \text{ Latm/mol K})(290 \text{ K})}{1.02 \text{ atm}} = 56.0 \text{ L H}_2\text{O}}$$

$$\text{Percent yield} = \frac{\text{actual mass}}{\text{theoretical mas}} \text{ x } 100\% = \frac{64.0 \text{ g}}{79.2 \text{ g}} \text{ x } 100\% = 80.8\%$$

(for CO_2)

$$\text{actual mass (for water)} = (0.808)(43.2 \text{ g}) = 39.4 \text{ g H}_2\text{O}$$

(10) Solutions

Solution: a homogeneous mixture.
Solute: substance that is dissolved.
Solvent: substance that does the dissolving (often water).
ex. in a solution of salt water, salt is the solute and water is the solvent.
Concentration: the amount of solute in a given amount of solution.
Molarity is another word for concentration.

Calculating Concentration/Molarity

Concentration can be calculated according to the following equation:

$$C = \frac{n}{V}$$

C = concentration (M)
n = number of moles (mol)
V = volume of solution (L)

Concentration is expressed in units of mol/L or Molar (M).
Brackets can be used to express the concentration of a chemical, i.e. [NaCl] is read as the "concentration of sodium chloride".
ex. What is the molarity of a 2.0 L solution that contains 0.026 mol of NaCl?

$$C = \frac{n}{V} = \frac{0.026\ \text{mol}}{2.0\ \text{L}} = 0.013\ \text{M}$$

ex. What is the volume of a 0.022 M solution that contains 0.11 mol of HCl?

$$V = \frac{n}{c} = \frac{0.11\ mol}{0.022\ M} = 5.0\ L$$

ex. How many grams of $ZnCl_2$ does 650 mL of a 2.40 M solution contain? How many molecules are present?

$$650\ \text{mL x} \frac{1\ \text{L}}{1000\ \text{mL}} 0.650\ \text{L}$$

$$\mathbf{n = CV = (2.40\ M)(0.650\ L) = 1.56\ mol}$$

$$\mathbf{1.56\ mol\ x\ \frac{136.29\ g}{mol} = 213\ g\ ZnCl_2}$$

$$1.56\ \text{mol x}\ \frac{6.02\text{x}10^{23}\ \text{molecules}}{\text{mol}} = 9.39\text{x}10^{23}\ \text{molecules}\ ZnCl_2$$

Solution Dilution

Solutions of a certain concentration are often prepared by diluting solutions of a higher concentration
Dilution: decreasing the concentration of a solution by adding more solvent (i.e. adding more water)

Solution Dilution Calculations

$$\mathbf{C_1V_1 = C_2V_2}$$

where: C_1 = initial concentration (M) C_2 = final concentration (M) (note: C_2 is always less than C_1)
V_1 = initial volume (L) V_2 = final volume (L) (note: V_2 is always more than V_1)

ex. 2.0 L of 0.24 M solution of KBr is diluted to a final volume of 4.8 L. Calculate the final concentration of the solution. Calculate the mass of KBr present in the solution.

$$\mathbf{C_1V_1 = C_2V_2}$$
$$\mathbf{(0.24\ M)(2.0\ L) = C_2(4.8\ L)}$$
$$\boldsymbol{C_2 = 0.10\ M}$$

$$\boldsymbol{n}$$
$$\boldsymbol{= CV\ (can\ use\ either\ C_1 and\ V_1\ OR\ C_2\ and\ V_2\ )}$$
$$\boldsymbol{= (0.24\ M)(2.0\ L)\ OR\ (0.10\ M)(4.8\ L)}$$
$$= 0.48\ \text{mol}$$

$$\boldsymbol{0.48\ mol\ x\ \frac{119\ g}{mol} = 57\ g}$$

ex. 125 mL of a solution contains 10.6 g of Na_2CO_3. What is the initial concentration of the solution? What is the final concentration if 275 mL of water are added to the solution?

$$\boldsymbol{125\ mL\ x\ \frac{1\ L}{1000\ mL} = 0.125\ L}$$

$$\boldsymbol{10.6\ g\ x\ \frac{mol}{105.99\ g} = 0.100\ mol\ Na_2CO_3}$$

$$C = \frac{n}{V} = \frac{0.100\ \text{mol}}{0.125\ \text{L}} = 0.800\ \text{M}\quad \text{INITIAL CONCENTRATION, } C_1$$

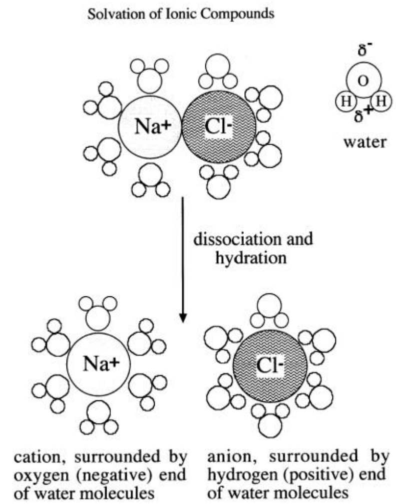
$$\boldsymbol{275\ mL\ x\ \frac{1\ L}{1000\ mL} = 0.275\ L\quad V_2 = 0.125\ L + 0.275\ L = 0.400\ L}$$

$$\mathbf{C_1V_1 = C_2V_2}$$
$$\mathbf{(0.800\ M)(0.125\ L) = C_2(0.400\ L)}$$
$$\boldsymbol{C_2 = 0.250\ M}$$

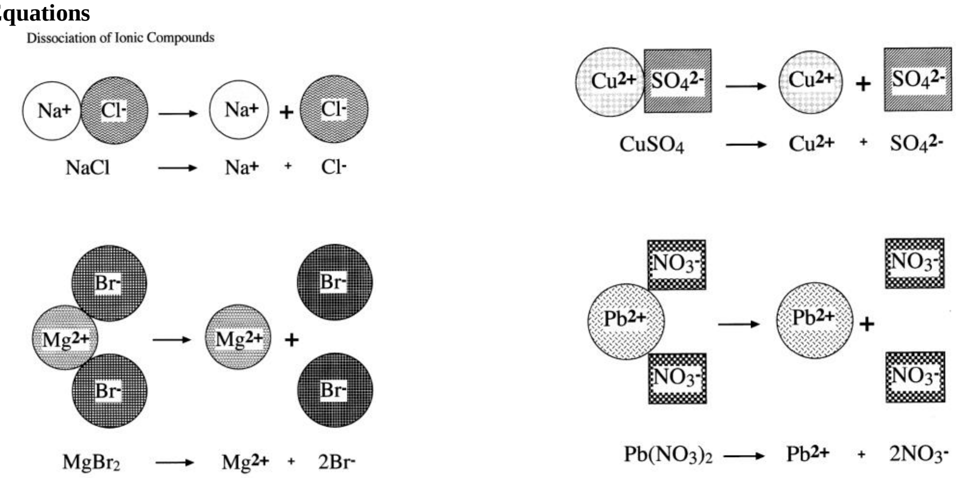
Solvation of Ionic Compounds

S**olvation**: the interactions between the solute and water.

Ionic compounds dissolve in water. As an ionic compound dissolves in water, the positive ion (cation) of the ionic compound interacts with the oxygen end of the water molecule (negative end) and the negative ion (anion) of the ionic compound interacts with the hydrogen end of the water molecule (positive end). These interactions result in **dissociation**: the separation of an ionic compound into its constituent ions. The ions then undergo **hydration**: they become surrounded by water molecules.



Dissociation Equations



Notice that polyatomic ions remain together

Write the dissociation equation for each compound.

(1) $BaBr_2$
$BaBr_2 \rightarrow Ba^{2+} + 2Br^{-}$

(2) $ZnSO_4$
$ZnSO_4 \rightarrow Zn^{2+} + SO_4^{2-}$

(3) cobalt (III) chloride
$CoCl_3 \rightarrow Co^{3+} + 3Cl^{-}$

(4) scandium chlorate
$Sc(ClO_3)_3 \rightarrow Sc^{3+} + 3ClO_3^{-}$

Dissociation and Ion Concentration

A dissociation equation is first written for the compound. To find the concentration of each ion, the concentration of the solution is multiplied by the coefficient in the dissociation equation of each ion. The concentration of each ion is listed.

ex. Calculate the concentration of each ion in a 0.26 M solution of aluminum nitrate

$Al(NO_3)_3 \rightarrow Al^{3+} \quad + \quad 3NO_3^{-}$

0.26 M   1(0.26 M)   3(0.26 M)
    = 0.26 M   = 0.78 M

$[Al^{3+}]$ = 0.26 M $[NO_3^{-}]$ = 0.78 M

ex. Calculate the concentration of each ion in a 0.0540 M solution of iron (III) sulphate

$Fe_2(SO_4)_3 \rightarrow 2Fe^{3+} \quad + \quad 3SO_4^{2-}$

0.0540 M   2(0.0540 M)   3(0.0540 M)
    = 0.108 M   = 0.162 M

$[Fe^{3+}]$ = 0.108 M $[SO_4^{2-}]$ = 0.162 M

Solubility
Solubility: the amount of solute that will dissolve in a solvent at a given temperature to form a **saturated** solution.
Saturated: a solution in which no more solute can be dissolved.
Unsaturated: a solution in which more solute can be dissolved.
The solubility for different solutes in water varies- for example the solubility of NaCl in water (at 25 ºC) is 6.0 M (about 350 g NaCl dissolve in one liter) whereas the solubility of AgCl in water (at 25 ºC) is 0.000010 M (about 0.0014 g AgCl dissolve in one liter).
Compounds that have high solubility are said to be **soluble** in water; compounds that have low solubility are said to be **insoluble** in water.
Soluble: more than 0.10 M (more than 0.10 moles of a solute dissolves in 1 L of water).
Insoluble: less than 0.10 M (less than 0.10 moles of a solute dissolves in 1 L of water).
Note: In general, the solubility of a compound increases with increased temperature.

Solubility Rules:
Soluble or mostly soluble
(1) All nitrate (NO_3^-) compounds are **soluble**
(2) All compounds containing alkali metal ions (Li^+, Na^+, K^+, Cs^+, and Rb^+) or ammonium ions (NH_4^+) are **soluble**
(3) Most chloride (Cl^-), bromide (Br^-), and iodide (I^-) compounds are **soluble** EXCEPT those containing the ions Ag^+, Pb^{2+}, Cu^+, and Hg_2^{2+}.
(4) Most fluoride (F^-) compounds are **soluble** EXCEPT those containing the ions Ca^{2+}, Ba^{2+}, Sr^{2+}, Mg^{2+}, and Pb^{2+}
(5) Most sulphate (SO_4^{2-}) compounds are **soluble** EXCEPT those containing the ions Ca^{2+}, Ba^{2+}, Sr^{2+}, Ag^+, and Pb^{2+}.

Mostly insoluble
(6) Most hydroxide (OH^-) compounds are **insoluble** EXCEPT those containing the ions Ba^{2+} and Sr^{2+}. (Ca^{2+} is sparingly soluble)
(7) Most sulphide (S^{2-}) compounds are **insoluble** EXCEPT those containing the ions Be^{2+}, Mg^{2+}, Ca^{2+}, Ba^{2+}, and Sr^{2+}.
(8) Almost all carbonate (CO_3^{2-}), phosphate (PO_4^{3-}), and chromate (CrO_4^{2-}) compounds are **insoluble**.

ex. Classify the following compounds as soluble or insoluble in water

(a) PbI_2	(b) MgS	(c) $Al(OH)_3$	(d) Li_2CO_3	(e) MgF_2	(f) $PbSO_4$	(g) $CuCl_2$
Answers:						
insoluble	soluble	insoluble	soluble	insoluble	insoluble	soluble (Cl^- is insoluble with Cu^+, not Cu^{2+}) CuCl would be insoluble

Solution Stoichiometry
ex. 100. mL of 0.300 M barium chloride reacts with 200. mL of sodium sulphate.
(a) What concentration of sodium sulphate solution is required?
(b) What would be the mass of each of the products?

	$BaCl_2$ +	Na_2SO_4 →	$BaSO_4$ +	$2NaCl$
moles (mol)	0.0300	0.0300	0.0300	0.0600
mass (g)			7.00	3.51
volume (L)	**0.100**	**0.200**		
Molarity (M)	**0.300**	0.150		
molar mass (g/mol)			137.33 + 32.06 + 4(16.00) =233.39	22.99 + 35.45 =58.44

$$100.\ mL\ x \frac{1L}{1000\,mL} = 0.100\,L \quad 200.\ mL\ x \frac{1L}{1000\,mL} = 0.200\,L$$

$$n = CV = (0.300\,M)(0.100\,L) = 0.0300\,mol\ BaCl_2$$

$$0.0300\,mol\ BaCl_2\ x \frac{1\,mol Na_2SO_4}{1\,mol\ BaCl_2} = 0.0300\,mol\ Na_2SO_4$$

$$C = \frac{n}{V} = \frac{0.0300\,mol}{0.200\,L} = 0.150\,M\ Na_2SO_4$$

$$0.0300\,mol\ BaCl_2\ x \frac{1\,mol BaSO_4}{1\,mol\ BaCl_2} = 0.0300\,mol\ BaSO_4 \quad 0.0300\,mol\ x \frac{233.39g}{mol} = 7.00\,g\ BaSO_4$$

$$0.0300\,mol\ BaCl_2\ x \frac{2\,mol NaCl}{1\,mol\ BaCl_2} = 0.0600\,mol\ NaCl \quad 0.0600\,mol\ x \frac{58.44\,g}{mol} = 3.51\,g\ NaCl$$

ex. 150. mL of 0.150 M lead (II) nitrate reacts with 200. mL of 0.250 M potassium iodide solution.
(a) Which reactant is limiting and which is excess?
(b) What would be the mass of each of the products?

	$Pb(NO_3)_2$ +	$2KI$ →	PbI_2 +	$2KNO_3$
moles (mol)	0.0225	0.0500	0.0225	0.0450
mass (g)			10.4	4.55
volume (L)	**0.150**	**0.200**		
Molarity (M)	**0.150**	**0.250**		
molar mass (g/mol)			207.2 + 2(126.91) 461.02	39.10 + 14.01+ 3(16.00) = 101.11

$$150.\ mL\ x \frac{1L}{1000\,mL} = 0.150\,L \quad 200.\ mL\ x \frac{1L}{1000\,mL} = 0.200\,L$$

$$n = CV$$

$$n = (0.150\,M)(0.150\,L) \qquad n = (0.250\,M)(0.200\,L)$$

$$n = 0.0225\,mol\ Pb(NO_3)_2 \qquad n = 0.0500\,mol\ KI$$

$$0.0225/1 = 0.0225 \qquad 0.0500/2 = 0.0250$$

$$Pb(NO_3)_2\,is\,limiting \qquad KI\ is\,excess$$

$$0.0225\,mol\ Pb(NO_3)_2\ x \frac{1molPbI_2}{1mol\ Pb(NO_3)_2} = 0.0225\,mol\ PbI_2 \quad 0.0225\,mol\ x \frac{461.02\,g}{mol} = 10.4\,g\ PbI_2$$

$$0.0225\,mol\ Pb(NO_3)_2\ x \frac{2\,mol\ KNO_3}{1mol\ Pb(NO_3)_2} = 0.0450\,mol\ KNO_3 \quad 0.0450\,mol\ x \frac{101.11\,g}{mol} = 4.55\,g\ KNO_3$$

(c) Calculate the concentration of nitrate ion, $[NO_3^-]$ at the *completion of the reaction*. Assume that volumes are additive.

At the end of the reaction, there are 0.0450 mol KNO_3. Since 0.150 L was added to 0.200 L of solution, the total volume of solution at the completion of the reaction is 0.350 L. The concentration of KNO_3 can be found from the moles and the volume.

$$\mathrm{C} = \frac{\mathrm{n}}{\mathrm{v}}$$

$$[\mathrm{KNO_3}] = \frac{0.0450\ \mathrm{mol}}{0.350\ L} = 0.129\ M$$

The concentration of the nitrate ion can be determined from the dissociation equation:

$KNO_3 \rightarrow K^+ + NO_3^-$
0.129 M 0.129 M

$[NO_3^-]$ = 0.129 M

Gravimetric Analysis
In gravimetric analysis, a substance is added to a solution that reacts specifically with a dissolved analyte to form an insoluble solid or **precipitate**. The mass of the solid formed can be used to calculate the amount of the analyte in the original solution.

ex. A 0.5660 g sample of a mineral containing chloride ions (as well as other elements) is dissolved in water. The solution is then reacted with silver nitrate, forming 1.010 g of silver chloride precipitate. What is the percent by mass of chlorine in the sample of the mineral?

$$1.010\ \mathrm{g\ AgCl\ x}\ \frac{\mathrm{mol\ AgCl}}{143.32\ \mathrm{g\ AgCl}}\ \mathrm{x}\ \frac{1\ \mathrm{mol\ Cl}}{1\ \mathrm{mol\ AgCl}}\ \mathrm{x}\ \frac{35.45\ \mathrm{g\ Cl}}{1\ \mathrm{mol\ Cl}} = 0.2498\ \mathrm{g\ Cl}$$

$$Percent\ by\ mass = \frac{mass\ element}{mass\ sample}\ x\ 100\% = \frac{0.2498\ \mathrm{g\ Cl}}{0.5660\ \mathrm{g\ sample}}\ x\ 100\% = 44.14\%\ Cl$$