UNIT 0 : AP Chemistry

TOPIC 0A: Chemistry, Scientific Method and Chemical & Physical Change

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).
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 and 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 nor destroyed).

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, respectively.

Physical change

If some aspect of the physical state of matter is altered, but the chemical composition remains the same, the change is a physical change. The most common physical changes are changes of state. These 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 move 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 or 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.

TOPIC 0B: Measurement

Measurements

Measurements, and subsequently calculations applied to those measurements, allow the determination of some of the quantitative properties of a substance; for example, mass and density.

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

For example, the number 42000.0 is converted to 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.

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

Task 0Bi

1 Convert the following numbers to scientific notation.

- ***(a) 24500***
- ***(b) 356***
- ***(c) 0.000985***
- ***(d) 0.222***
- ***(e) 12200***

2. Convert the following scientific notation numbers 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 scale that is being used for measurement. Prefixes are used to make writing very large or small numbers easier. 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}

Converting units and 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 of 1.00 inch = 2.54 cm, the conversion factor will either be,

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

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

$$\frac{9.00 \text{ in}}{} \left| \frac{2.54 \text{ cm}}{1.00 \text{ in}} \right.$$

To convert 5.00 cm into inches, perform the following calculation

$$\frac{5.00 \text{ cm}}{} \quad \frac{1.00 \text{ in}}{2.54 \text{ cm}}$$

Task 0Bii

1. ***Convert the following quantities from one unit to another using the equivalence statements.***
   1.000 m = 1.094 yd, 1.000 mile = 1760 yd, 1.000 kg = 2.205 lbs

   (a) 30 m to miles
   (b) 1500 yd to miles
   (c) 206 miles to m
   (d) 34 kg to lbs
   (e) 34 lb to kg

2. ***Which is the larger quantity in each case below?***

   (a) ***A distance of 3.00 miles or 3000. m.***

   (b) ***A mass of 10.0 kg or 25 lbs.***

Temperature

There are three scales of temperature that you may come across in your study of chemistry. They are Celsius (ºC), Fahrenheit (ºF) and Kelvin (K). The following conversion factors will be useful.

Temperature Conversion factors	
Celsius to Kelvin	$\boxed{K = C + 273^{o}}$
Kelvin to Celsius	$\boxed{C = K - 273^{o}}$
Celsius to Fahrenheit	$\boxed{F = (1.8 \times C) + 32^{o}}$
Fahrenheit to Celsius	$\boxed{C = \dfrac{(F^{\circ} - 32)}{1.8}}$

Task 0Biii

1. ***Convert the following temperatures from one unit to the other.***

   (a) ***263 K to ºF***

   (b) ***38 K to ºF***

   (c) ***13 ºF to ºC***

   (d) ***1390 ºC to K***

   (e) ***3000 ºC to ºF***

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 in chemistry is volume. Volume has the unit, $length^3$. Common units for volume are liters (L) or milliliters (mL).

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

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

Density is the ratio of the mass to volume.

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

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.

$$\text{mass} = (\text{density})\ (\text{volume}) = (\text{g/L})(\text{L})$$

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.

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 number of significant figures present in a number

1. Any non-zero integers are always counted as significant figures.
2. Leading zeros are those that precede all of the 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 an unlimited number of significant figures. (Exact numbers are those which are as a result of counting e.g., 3 apples or 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.

Determining the correct number of significant figures to be shown as the result of a calculation

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 decimal 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.

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

Task 0Biv

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) (2.61×10^{-1}) ***(356)***
   (d) 21.78 + 45.86
   (e) 23.888897 - 11.2
   (f) 6 - 3.0

Accuracy and 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.

Task 0Bv

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?

Percentage 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 *percentage error*.

$$\text{Percentage Error} = \frac{(\text{Actual Value - Calculated Value})}{\text{Actual Value}} \text{ x } 100$$

TOPIC 0C: 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

LO 1.13

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 new compounds. (Law of conservation of mass).

Structure of the atom and the periodic table

Several experiments were being carried out in the 19th and 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 on the electron	Electron
Rutherford, Marsden and 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*. Schrödinger subsequently expanded upon Bohr's model, in order to incorporate the wave nature of 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	$\frac{1}{1836}$	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 electrons surrounding the nucleus.

N.B., when atoms lose or gain electrons the proton and 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 c*ation*.

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

Period	GROUP 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H 1																	2 He 4
2	3 Li 7	4 Be 9											5 B 11	6 C 12	7 N 14	8 O 16	9 F 19	10 Ne 20
3	11 Na 23	12 Mg 24	TRANSITION METALS										13 Al 27	14 Si 28	15 P 31	16 S 32	17 Cl 35.5	18 Ar 36
4	19 K 39	20 Ca 40	21 Sc 45	22 Ti 48	23 V 51	24 Cr 52	25 Mn 55	26 Fe 56	27 Co 59	28 Ni 59	29 Cu 64	30 Zn 65	31 Ga 70	32 Ge 73	33 As 75	34 Se 79	35 Br 80	36 Kr 84
5	37 Rb 86	38 Sr 88	39 Y 89	40 Zr 91	41 Nb 93	42 Mo 96	43 Tc 99	44 Ru 101	45 Rh 103	46 Pd 106	47 Ag 108	48 Cd 112	49 In 115	50 Sn 119	51 Sb 122	52 Te 128	53 I 127	54 Xe 131
6	55 Cs 133	56 Ba 137	57 La* 139	72 Hf 178	73 Ta 181	74 W 184	75 Re 186	76 Os 190	77 Ir 192	78 Pt 195	79 Au 197	80 Hg 201	81 Tl 204	82 Pb 207	83 Bi 209	84 Po 210	85 At 210	86 Rn 222
7	87 Fr 223	88 Ra 226	89 Ac† 226	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Fl	115 Uup	116 Lv	117 Uus	118 Uuo

*Lanthanides	58 Ce 140	59 Pr 141	60 Nd 144	61 Pm 147	62 Sm 150	63 Eu 152	64 Gd 157	65 Tb 159	66 Dy 163	67 Ho 165	68 Er 167	69 Tm 169	70 Yb 173	71 Lu 175
†Actinides	90 Th 232	91 Pa 231	92 U 238	93 Np 237	94 Pu 242	95 Am 243	96 Cm 247	97 Bk 251	98 Cf 251	99 Es 254	100 Fm 253	101 Md 256	102 No 254	103 Lr 257

KEY:

Metal	Semi Metal	Non-metal

13 Al 27	14 Si 28	15 P 31

In this example Al is a metal, Si is a semi-metal (metalloid) and P is a non-metal.

Task 0Ci

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 numbers of protons and neutrons

TOPIC 0D: Nomenclature

Nomenclature

Nomenclature is the language of chemistry, and a grasp of it is essential to studying the subject.

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 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 example, Co (cobalt), a metal element, 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 compounds)

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 a 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.

A few common ions, their charges and formulae are listed below. You will need a more complete list, found here.

Negative ions (ANIONS)			Positive ions (CATIONS)		
Name	**Charge**	**Symbol**	**Name**	**Charge**	**Symbol**
Bromide	1-	Br^-	Aluminum	3+	Al^{3+}
Chloride	1-	Cl^-	Barium	2+	Ba^{2+}
Fluoride	1-	F^-	Calcium	2+	Ca^{2+}
Hydride	1-	H^-	Copper (I)	1+	Cu^+
Iodide	1-	I^-	Copper (II)	2+	Cu^{2+}
Nitride	3-	N^{3-}	Hydrogen	1+	H^+
Oxide	2-	O^{2-}	Iron (II)	2+	Fe^{2+}
Phosphide	3-	P^{3-}	Iron (III)	3+	Fe^{3+}
Sulfide	2-	S^{2-}	Lead (II)	2+	Pb^{2+}
			Lead (IV)	4+	Pb^{4+}
			Lithium	1+	Li^+
			Magnesium	2+	Mg^{2+}
			Manganese (II)	2+	Mn^{2+}
			Nickel (II)	2+	Ni^{2+}
			Potassium	1+	K^+
			Silver	1+	Ag^+
			Sodium	1+	Na^+
			Strontium	2+	Sr^{2+}
			Tin (II)	2+	Sn^{2+}
			Tin (IV)	4+	Sn^{4+}
			Zinc	2+	Zn^{2+}

Most transition metal ions (and a few other metal ions) include a Roman 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.

Task 0Di

1. ***Name these binary compounds.***

   - ***(a) NaCl***
   - ***(b) SrO***
   - ***(c) AlN***
   - ***(d) $BaCl_2$***
   - ***(e) K_2O***
   - ***(f) CuO***
   - ***(g) Cu_2O***

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 other 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. You will need a more complete list, found here.

Common Polyatomic ions

Name	Charge	Formula
Ammonium	1+	NH_4
Carbonate	2-	CO_3
Chromate (VI)	2-	CrO_4
Dichromate (VI)	2-	Cr_2O_7
Ethanedioate	2-	C_2O_4
Hydrogen carbonate	1-	HCO_3
Hydrogen sulfate	1-	HSO_4
Hydroxide	1-	OH^-
Manganate (VII) (permanganate)	1-	MnO_4
Nitrate	1-	NO_3
Nitrite	1-	NO_2
Phosphate	3-	PO_4
Sulfate	2-	SO_4
Sulfite	2-	SO_3

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	Increase in number of oxygen atoms
*(element)*ite	
*(element)*ate	
Per*(element)*ate	

Where there are only two members in such a series the endings are –ite and –ate. 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

Task 0Dii

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) ***Aluminum sulfate***
   (f) ***Sodium perchlorate***
   (g) ***Copper (II) iodite***

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 compounds)

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!

Task 0Diii

1. ***Write formula or names for the following molecular compounds.***

    (a) Dinitrogen tetroxide

    (b) Phosphorous 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, $CuSO_4.5H_2O$ is copper (II) sulfate pentahydrate.