ATOMS, MOLECULES and IONS

[MH5; Ch. 2]

Atoms and Atomic Theory [MH5; 2.1 - 2.2]

- The three main postulates of atomic theory are:
  1) Elements are made up of atoms.
     - all atoms of an element - same prop.
     - a different element ? - different prop.
  2) In chemical reactions, atoms transfer from one substance to another but they do not disappear.
     - can't create or destroy atoms in a chemical reaction
  3) Compounds are formed when two or more elements combine in a chemical reaction.
     - any compound: number and kind of each atom is constant
- So, knowing that elements are the building blocks of everything else, we can say that the **ATOM** is the smallest entity that we deal with.
- Atoms are **neutral** species made up of a small positive nucleus, surrounded by a number of negative electrons (abbreviated e^-).
- The number of electrons is equal to the positive charge on the nucleus.
- The nucleus accounts for most of the atom's mass and contains two different types of particles: protons (+1 charge) and neutrons (have no charge).
- It follows that in a neutral atom, the number of electrons outside the nucleus is equal to the number of protons in the nucleus.

EXAMPLES: He atom: He^{2+} (in nucleus) + 2 e^- (outside nucleus)
Ca atom: Ca^{20+} (in nucleus) + 20 e^- (outside nucleus)

- The positive charge (the number of protons) on the nucleus is the **Atomic Number** of that atom, **Z**. (This is the number above the element on the Periodic Table.)
- A sample of any given **ELEMENT** is comprised of many atoms, all of which have the same atomic number; in Carbon, every atom has **6** e^- and a nucleus with a charge of **+ 6** (ie; 6 protons).
- The atoms of each element have a unique number of protons, neutrons and electrons; it is these differing numbers that make the elements behave differently.
- **ISOTOPES** are atoms of the same element (ie, same atomic number) but the atoms differ in mass; due to differences in the total number of particles in the **nucleus**.
- Isotopes of an element have different **Mass Numbers**, denoted by ***A***.

EXAMPLE: Carbon has three (3) isotopes

Mass Number (***A***): 12, 13, 14
Atomic Number (**Z**): 6, 6, 6

$^{12}_{6}C$, $^{13}_{6}C$, $^{14}_{6}C$

Mass Number (*A*) = # of protons + # of neutrons
= **Atomic Number (Z)** + # of neutrons **OR...**

Mass Number (*A*) - **Atomic Number (Z)** = # of neutrons

- In $^{12}_{6}C$: **12** (A) - **6** (Z) = 6 neutrons

The Periodic Table [MH5; 2.3]

- The Periodic Table is a listing of all known elements in Atomic Number order.
- Recall that an element consists of atoms which have the same number of protons, and therefore, the same Atomic Number.
- Chemical properties of elements depend on the atomic number of the element.
- A complete Periodic Table lists the elements, their symbols and atomic numbers as well as atomic masses.

- The Periodic Table is arranged into rows, called **periods** and columns, which are called **groups**. → sometimes called Rows
- The first period consists of only hydrogen (H) and helium (He); the second period starts at lithium (Li) and ends at neon (Ne).
- The groups are numbered from 1 (for the column headed by H) to 18 (for the column headed by He).
- Elements in Groups 1, 2, 13, 14, 15, 16, 17 and 18 are called the **Main Group** elements; those in the centre (Groups 3 to 12) are called the **Transition Metals**.
- Elements in Groups 13, 14 and 15 are sometimes termed **Post Transition Metals.**
- Some main group elements have group names.....Group 1 are the **alkali metals** and Group 2 are the **alkaline earth metals**.
- Group 17 elements are known as the **halogens** and Group 18 elements are called **noble**, or **inert** (ie; unreactive) gases.
- Elements belonging to a certain group all exhibit similar chemical properties.

Molecules and Ions [MH5; 2.4]

- Isolated atoms are not commonly found in nature; it is only the inert gases (Group 18) that exist as non reactive atoms.
- All other elements will combine with each other (lots of possibilities!) to form more complex units which are called **molecules**.
- A **molecule** is two or more atoms joined together by **covalent bonds**; formed by the sharing of electrons by the atoms involved.
- The atoms forming a molecule may be of the same element, but are usually of different elements.
- A **compound** is a substance in which atoms of more than one element are present, usually in an integral (whole number) ratio.
- A chemical combination has occurred to create this compound.

EXAMPLES:

H_2O 1 molecule of water:

2 Hydrogen atoms

1 Oxygen atom

NH_3 1 molecule of NH_3 (ammonia):

1 Nitrogen atom

3 Hydrogen atoms

CH_4 1 molecule of CH_4 (methane):

1 Carbon atom

4 Hydrogen atoms

- Chemical combination may occur between atoms of the same element, but by convention the substance is still called an element, not a compound.
- The **molecular elements** are:

H_2 (g), N_2 (g), O_2 (g), F_2 (g), $C\ell_2$ (g), Br_2 (ℓ), I_2 (s), P_4 (s), S_8 (s)

Ions

- Ions are **charged** species.
- In a **simple ion**, one nucleus is present, but the species carries a charge because the number of electrons does not equal the +ve charge on the nucleus.
- This means that the atom has either lost or gained one or more electrons........
- A gain of electrons results in a negatively charged ion; known as an **ANION**. *Negative*
- A loss of electrons results in a positively charge ion; known as a **CATION**. *Cation + in positive*

EXAMPLES:

F^{-} (ANION) = ${}_{9}F^{9+}$ (nucleus), 10 e^{-}

9 protons ⇒ +9
10 electrons ⇒ −10
−1

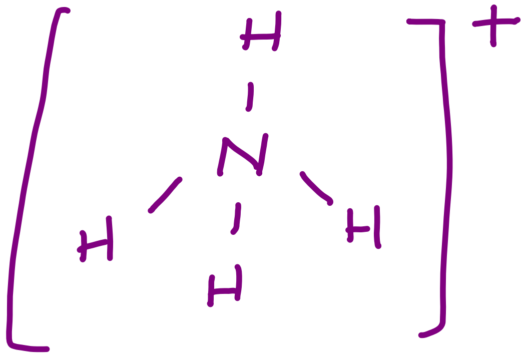
Ca^{2+} (CATION) = ${}_{20}Ca^{20+}$ (nucleus), 18 e^{-}

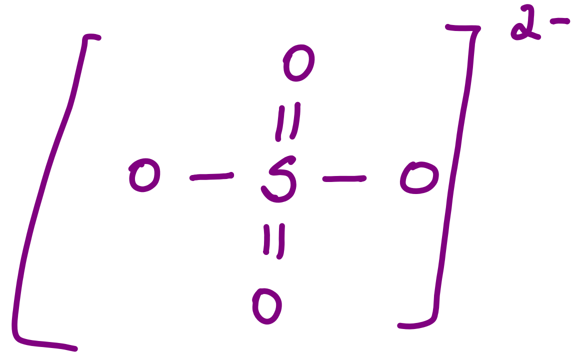
20 protons ⇒ +20
18 electrons ⇒ −18
+2

- Notice that the addition of one electron to Fluorine (and also the removal of two electrons from Calcium) result in these ions having the same number of electrons as Neon (and Argon).
- This results in a very **stable** species; most **main group elements** will try to achieve this number of electrons when forming ions.

- A **COMPLEX ION** is like a molecule, several atoms joined by covalent bonds, but there is an overall charge.

EXAMPLES: NH_4^+ SO_4^{2-}





- Compounds can also be formed by the interaction of anions (negatively charged species) and cations (positively charged species).
- These are called **ionic compounds**, and although they contain charged species, they are electrically neutral......the number of positive charges will equal the number of negative charges.
- These compounds are held together by electrostatic attraction between the positive and negative ions; these forces are called **ionic bonds**. Sodium chloride : NaCl

$Na^{\oplus}$ ------ $Cl^{\ominus}$

- Ionic compounds are usually solids at room temperature and have high melting points.
- When ionic compounds dissolve in water, a solution is formed which contains both anions and cations.......

$$NaCl(s) \xrightarrow{H_2O} \underset{\text{Cation}}{Na^+(aq)} + \underset{\text{Anion}}{Cl^-(aq)}$$

- This solution will conduct electricity and the solute is termed an **electrolyte**.....
- A solution in which the solute is a molecular solid does not conduct electricity; the solute is called a **nonelectrolyte**. ↳ doesn't form ions

Formulas of Ionic Compounds [MH5; 2.5]

- The formula of an ionic compound shows the simplest whole number ratio of cations to anions; the metal is always shown first.
- If you know the charges of the two ions involved, it is easy to predict the formula of the ionic compound.
- The number of positive charges must equal the number of negative charges. Ionic Cmpds: a metal from Gr 1 or 2 and an element from RHS of Periodic Table

EXAMPLES:

element → Ion → Ionic Cmpds

Na and Cl: $Na \rightarrow Na^{+}$, $Cl \rightarrow Cl^{-}$

$Na^{+} + Cl^{-} \rightarrow NaCl$

Mg and F: $Mg \rightarrow Mg^{2+}$, $F \rightarrow F^{-}$

$Mg^{2+} + 2F^{-} \rightarrow MgF_2$

- How does one predict the charges on an ion ??
- Look at the Periodic Table; atoms that are close to an inert gas (Group 18) will either gain or lose electrons to attain the same number of electrons as the inert gas.

EXAMPLES:

O: #8, close to Ne (#10)

$\therefore O \Rightarrow O^{2-}$

Ba: #56, close to Xe (#54)

$\therefore Ba \Rightarrow Ba^{2+}$

- Many of the transition metals and post transition metals form more than one cation; they do not normally have an inert gas structure.
  - figure out charge based on the other element

EXAMPLES:

IRON (Fe)

$FeBr_2 \Rightarrow Fe^{2+}$

$FeBr_3 \Rightarrow Fe^{3+}$

COPPER (Cu)

$CuCl \Rightarrow Cu^{+}$

$CuCl_2 \Rightarrow Cu^{2+}$

* Both Br and Cl are halogens ⇒ form -1 ions

- There are many polyatomic, or complex ions.....

  Ammonium ; NH_4^{+}

  Sulfate ; SO_4^{2-}

- Most polyatomic anions contain oxygen atoms; these species are known as **oxoanions**.

  Sulfate : SO_4^{2-}

  Nitrate : NO_3^{-}

  Carbonate : CO_3^{2-}

  Phosphate : PO_4^{3-}

  Oxoanions like to stay "stuck together"