What is Chemistry?

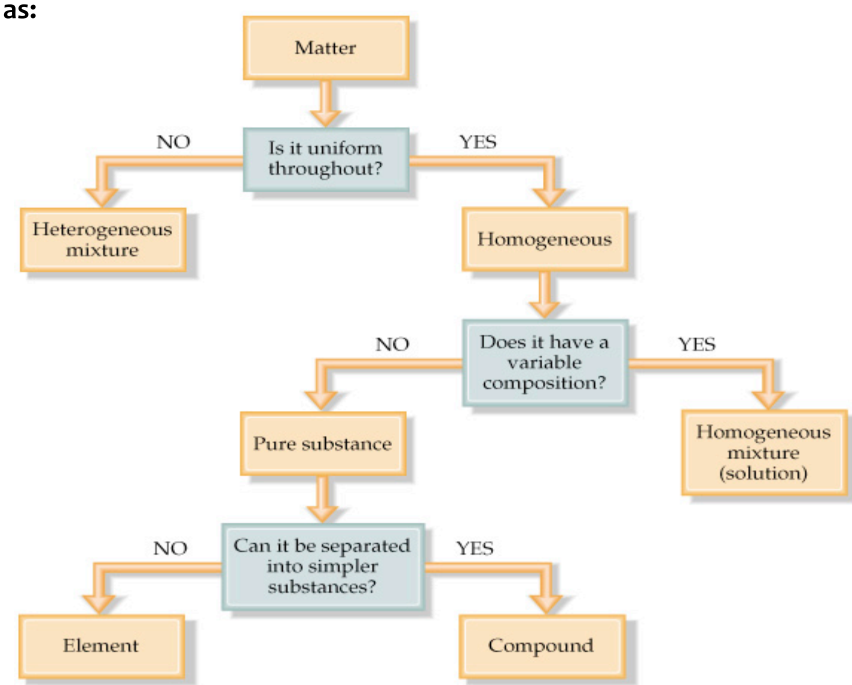
- **Chemistry** *is the physical science that deals with the composition, properties, and changes in matter.*
- Changes studied in chemistry include *chemical changes such as chemical reactions,* and *physical changes* such as ice freezing or the vaporization of a gas.
- Chemistry occurs *everywhere* around you and even *within* you each and every second of every day.

- Traditionally, there are five subdivisions used to classify various fields of study within chemistry:

Branch of Chemistry	***The study of…***
Organic Chemistry	compounds of hydrogen and carbon and their derivatives
Inorganic Chemistry	inorganic chemicals (do not contain carbon)
Biochemistry	chemicals and chemical change in living systems
Physical Chemistry	changes in the structure and energy of chemicals and chemical systems
Analytical Chemistry	separation, identification, and the quantification of matter

Unit 1.1 - Classifying Matter (Hebden p 45 & p 49 – 61)

- **Matter** is anything that has mass and occupies space → everything!!
- Types of matter are classified using an empirical (observable) classification system based on the methods used to separate matter.
- Common classification differentiates matter as pure substances and mixtures (impure substances)
- **Matter can be classified empirically as:**



Mixtures

- Most of the matter we encounter *consists of mixtures of different substances.*
- Mixtures can be physically separated by a variety of techniques including mechanical, settling, floatation, filtration, extraction, distillation, crystallization, chromatography, and using magnets to separate certain metals. * *See* Hebden p 53 - 59

1) **Heterogeneous mixtures** have variable composition and variable appearance throughout and may consist of more than one phase *(solid and/or, liquid, and/or gas)*

e.g. sand, rocks, and wood

2) **Homogeneous mixtures** are uniform throughout and consist of only one phase (*either – solid, liquid, or gas)* e.g. air, salt water

- Many substances dissolve in water to form homogeneous mixtures e.g. salt water, syrup, dye, etc.

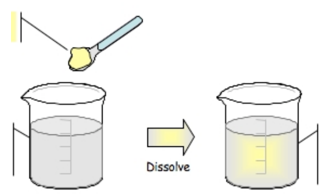
** Some substances that appear homogenous may, on closer inspection, prove to be heterogeneous.*

3) **Solutions** - homogeneous mixtures are *also* called **solutions.**

- Solutions contain particles of more than one substance (**solute** and **solvent** particles), uniformly distributed throughout them.

i) **Solute** – the quantity of *LESSER* amount in a solution.
- Usually the solute is a solid
e.g. salt in a saline solution

ii) **Solvent** – the quantity of *GREATER* amount in a solution.
- Usually the solvent is a liquid
e.g. water in a saline solution

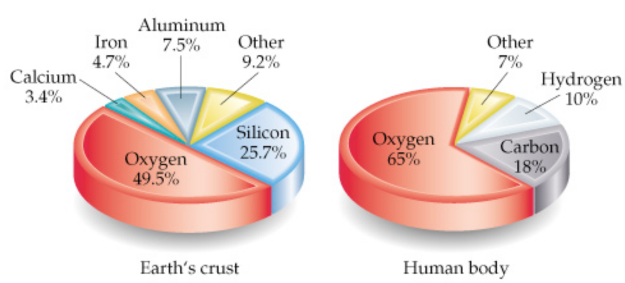


Pure Substances

1) **Elements** - composed of only one kind of atom and *CANNOT* be broken down into simpler chemical substances by any physical or chemical means.

- Elements are represented by a single symbol - *one or two* letters
- First letter capitalized, second letter is lower case

e.g.



i) **Atoms** - the smallest possible particle of an element that still retains the properties and characteristics of the element.

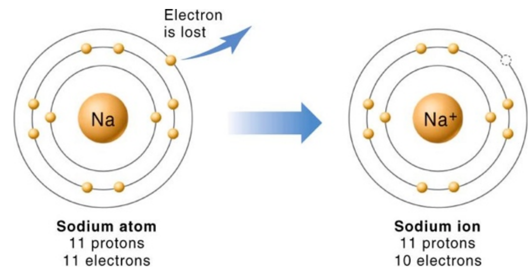
ii) **Ions** - charged particles (have gained or lost electrons).

- "+" indicates a loss of electrons

  e.g. Na^+ - a sodium atom has lost an electron

- "-" indicates a gain of electrons

  e.g. NO_3^- - nitrate molecule has gained an electron



2) **Compounds** – substances that contain atoms of more than one element combined in a definite, fixed proportion.

- Compounds are represented by chemical formulas that contain 2 or more different symbols.

  e.g. Water's chemical formula is H_2O - 2 hydrogen atoms and 1 oxygen atom make 1 water molecule

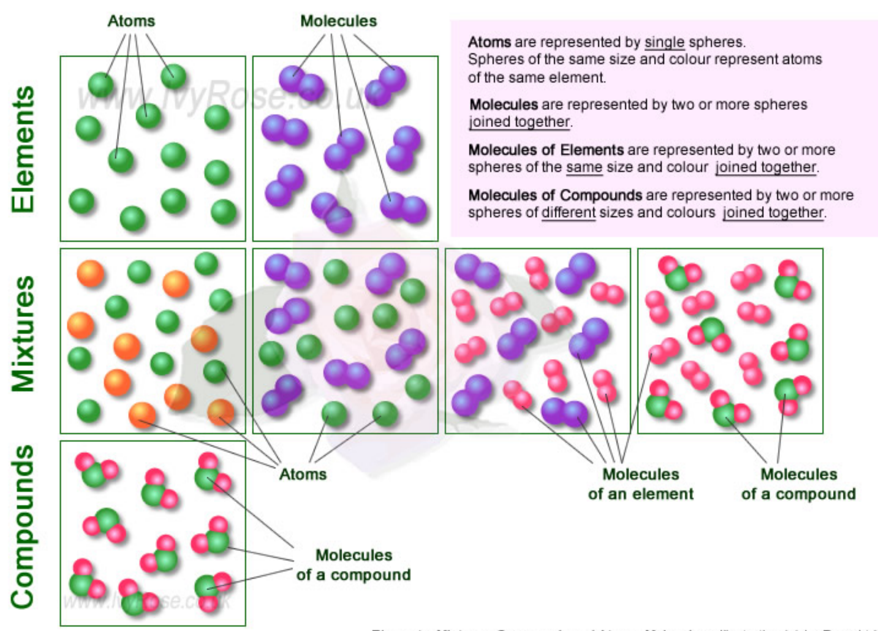
3) **Molecules** – substances that are composed of more than one atom.

- Can have molecules of an *element* – 2 or more atoms of the *SAME* element bound by chemical bonds to each other.

  e.g. O_2, Cl_2

- Can have molecules of a *compound* – 2 or more atoms of different elements bound by chemical bonds to each other.

  e.g. H_2O, NaOH, HCl



Elements, Mixtures, Compounds and Atoms, Molecules - Illustration (c) IvyRose Ltd. 2011.

Properties of Matter

- Properties are characteristics used to identify or describe a substance.
- May be *qualitative (descriptive*) or *quantitative (measurements, ALWAYS include a number)*

QUALITATIVE	***QUANTITATIVE***
State (s, l, g, aq)	*Solubility*

1) **Physical Properties of Matter**

- Properties of an element or compound that *can be observed without the chemical reaction of the substance.*

  e.g. density, electrical conductivity, melting point, and…

  - *Hardness* – ability of a solid to resist abrasions or scratching
  - *Malleability* – ability of a solid to be hammered into thin sheets
  - *Ductility* – is the ability to be stretched or drawn into wires
  - *Lustre* – the ability of a solid to reflect light – lustres can be metallic, glassy, oily, pearly, silky or dull.
  - *Viscosity* – fluid resistance to flow.

2) **Chemical Properties of Matter**

- Characteristics of a substance that *are observed when it undergoes a chemical change by itself or with other substances.*

  e.g. reactivity, flammability, chemical reaction where a new substance is created

States of Matter

- Matter exists in 3 common states or phases:

1) **Solid**

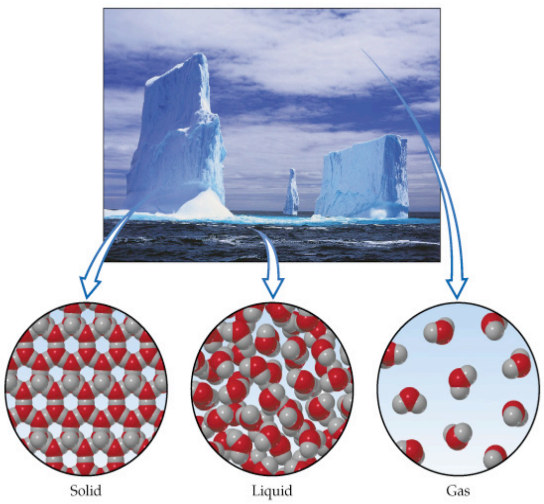
- Particles of solids are highly organized and are very close together.
- Close proximity and organized arrangement of solid particles causes solids to be rigid – they do not readily change their shape, and experience relatively small changes in volume when heated or subjected to pressure i.e. solids are not compressible.

2) **Liquid**

- Particles remain in close contact but have room to move past one another.
- The space between liquid particles allows liquids to conform to the shape of their container.
- Liquids experience only slight changes in volume with temperature changes or when subjected to pressure i.e. liquids are not compressible.
- Liquids share some properties with both solids and gases.

3) **Gas**

- Particles are widely separated and contact each other during collisions.
- Conform to shape of their container, experience *drastic* changes in volume when heated or subjected to pressure.
- Gases are compressible because the space between gas particles decreases when the volume of a container decreases or the pressure within a container increases.



- There are also several exotic states of matter including plasma, superconductive states, superfluid states, and supercondensed states (*for more info read Hebden p 46).*

Changes of Matter

1) **Physical Changes of Matter**

- Substances are *NOT* altered chemically, *NO* new products are formed, and chemical bonds are *NOT* broken – *NO* change in the written formula of the substance(s) involved.
- A physical change can affect the size, shape or color of a substance but *DOES NOT* affect its composition.
- Physical changes in matter usually involve relatively small amounts of energy change.
- Physical changes are *often reversible.*

  e.g. phase changes of substances - melting ice or boiling water (steam)
  $H_2O(s) \leftrightarrow H_2O\ (l) \leftrightarrow H_2O\ (g)$

e.g. breaking a glass, grinding a substance into a powder, or dissolving sugar in water

Phase Changes

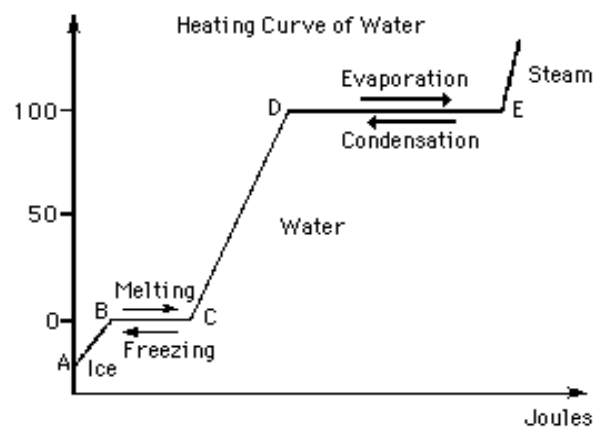
- A phase change is a change in the state of matter without any change in its chemical composition.
- Phase changes always involve energy changes but never involve temperature changes.
- Heating and cooling curves are used to show phase changes for matter:
- Common phase changes include:

  i) Solid ↔ Liquid

  ii) Liquid ↔ Gas

  iii) Solid ↔ Gas (called *sublimation* from solid → gas and *deposition* from gas → solid)

  - ***Melting point*** – the temperature at which a solid changes into the liquid phase.
  - ***Freezing point*** – the temperature at which a liquid changes into the solid phase.
  - ***Boiling point*** – the temperature at which a liquid changes into the gas phase.
  - ***Condensation point*** – temperature at which a gas changes into the liquid phase.



**** Level portions of a phase change graph*** – happen when a substance contains so much heat energy that it cannot absorb any more heat and stay in the same phase. All the added heat is used to facilitate phase changes i.e. solid → liquid or liquid → gas, this is why there is no temperature change during a phase change and the graph levels off.

2) **Chemical Changes of Matter**

- Substances are altered chemically and display different physical and chemical properties after the change (color, odour, or state).
- When a chemical change occurs at least one new substance is formed that has *different* physical and chemical properties than the original matter.
- Involve some kind of change in the chemical bonds within the atoms or ions of a substance.
- Involve larger energy changes than physical changes.
- Chemical changes are *irreversible*.

  e.g. rusting iron, frying an egg, burning a match

  e.g. decomposition of hydrogen peroxide into water and oxygen gas $2H_2O_2 \longrightarrow 2H_2O + O_2(g)$

3) **Nuclear Changes of Matter**

- Create entirely new atomic particles (*more on this later*).
- Represented by formulas that show new atomic symbols, different from those of the original matter.
- Nuclear changes involve *EXTREMELY LARGE* changes in energy and often emit radiation.

  e.g. Natural radioactive decay of uranium into thorium and helium. $^{238}_{92}U \longrightarrow ^{234}_{90}Th + ^{4}_{2}He$

  e.g. Nuclear fusion – conversion of hydrogen to helium (occurs on the surface of the sun).

  e.g. Nuclear fission – splitting the nucleus of atoms to release large amounts of energy (nuclear power plants).

Unit 1.2 – Elements & The Major Divisions of The Periodic Table

(Hebden p 160 – 164 & p 165 - p 178)

Classifying Elements – Chemical Families

- Since ancient times, people have known of seven metallic elements (gold, silver, copper, tin, lead, mercury, iron) – each was associated with a symbol of a different celestial body.
  e.g. gold - sun, copper – Venus, lead – Saturn, etc.
- Early 1800'*s* - alchemists had discovered new elements and needed new symbols to represent them.
- 1814, Swedish chemist Jons Jacob Berzelius used letters as symbols to represent elements.
- Because Latin was the common language of that time, many of the symbols on the modern periodic table today were derived from the Latin names for the elements.
- Today IUPAC (the *International Union of Pure and Applied Chemistry)* specifies the rules for chemical names and symbols of elements – IUPAC rules are used and are understood worldwide.
- All elements can be classified into one of 3 categories: metals, non-metals, or semiconductors (semiconductors were formerly known as metalloids).

Metals	***Non-Metals***	***Semiconductors (Metalloids)***
• Majority of all known elements are metals • Found on the left side of the "staircase" on the periodic table • Shiny • Bendable • Good conductors of heat and electricity • All metals (except mercury) are solid at STP • E.g. sodium (Na), iron (Fe)	• Found to the right side of the "staircase" on the periodic table • Not shiny • Not bendable • Poor conductors of heat and electricity in solid form • At STP most non-metals are gases, few are solids • Solid non-metals are brittle and lack the lustre of metals • E.g. Argon (Ar), oxygen (O)	• Elements that do not behave like metals or non-metals • Few of them • Found near the "staircase" on the periodic table • Very hard • Have high melting points • Are non-conductors or semiconductors of electricity • E.g. carbon (C) & silicon (Si)

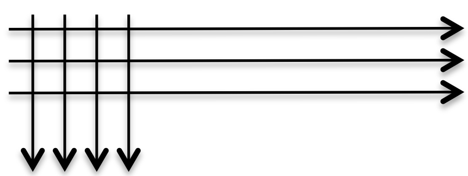
** STP – stands for "standard temperature and pressure" conditions.*
** STP refers to conditions where the temperature is exactly* $0°C$ $(273K)$ *and atmospheric pressure is exactly* $101.325\ kPa$ $(1\ atm)$.

- Some metals and non-metals are *ductile* – can be stretched into a wire or a tube.
- Some metals and non-metals are *malleable* – can be hammered into a thin sheet.
  e.g. Gold - one of the most malleable metals, can be hammered into a thin foil

Interactive photographic periodic table of elements that illustrates physical properties of metals, non-metals and metalloids: http://periodictable.com/

Major Divisions Within the Periodic Table

- Within the periodic table there are several special groups, rows and "blocks" of elements that share similar chemical and physical properties.
- *Horizontal rows* of elements going *across* the periodic table are called ***periods***.
  - As you move across the periodic table from left → right the properties of elements gradually become *less metallic* and *more non-metallic*.
  - There are 7 $periods$ on the modern periodic table.
- *Vertical columns* of elements are called ***groups, or families***.
  - There are 18 $groups$ on the modern periodic table.
  - **Hydrogen** – is a special element sometimes it behaves as a metal as part of $group\ 1$, sometimes it behaves like a non-metal.



1) **Metals** – to the left of the "staircase"

i) ***Alkali Metals*** - $Group\ 1$
- Soft, silver-coloured metals that react violently with water to form basic solutions
- All Alkali metals react with halogens ($group\ 17$) to form compounds similar to most salts
- All alkali metals are stored in oil or in a vacuum to prevent reaction with air

ii) ***Alkaline-Earth* Metals** - $Group\ 2$
- Light, reactive metals that form oxide coatings when exposed to air
- All alkaline-earth metals react with oxygen to form oxides
  e.g. MO, CaO, MgO

iii) ***Transition Metals*** - Range from $group\ 3\ to\ group\ 12$

- Exhibit a wide range of chemical and physical properties
- As the group number increases the transition metals exhibit characteristics more like non-metals and less like metals

2) **Non-metals** - to the right of the "staircase"

iv) ***Halogens*** - $Group\ 17$

- Extremely reactive, includes the element Fluorine, the most reactive element of all
- All halogens react with hydrogen to form hydrogen halides
- Water solutions of all halides are acidic
  e.g. HCl (hydrochloric acid), HI

v) ***Noble Gases*** - $Group\ 18$

- Extremely stable, extremely low chemical reactivity (once thought to be non-reactive)
- Are only elements on the periodic table with completely full electron shells

- In addition to the common classes of elements we described above, the bottom two rows (periods) in the periodic table each have a name as well.

vi) **Lanthanides**

- Rare-earth elements
- Have atomic numbers $58\ to\ 71$ (period begins with the element *lanthanum*)
- Are metallic chemical elements
- Used as catalysts and in the production of glasses

vii) **Actinides** -

- Have atomic numbers $90\ to\ 103$ (period begins with the element *actinium*)
- All actinides are radioactive and release energy when they radioactively decay
- Nuclear weapons release at least 6 of the actinides into the environment when detonated

viii) **Transuranic Elements**

- Elements that do not naturally occur in nature
- Elements that have atomic numbers of 93 or greater
- None are stable and all radioactively decay into other elements
- Those that can be found on Earth now are artificially generated synthetic elements, via nuclear reactors or particle accelerators

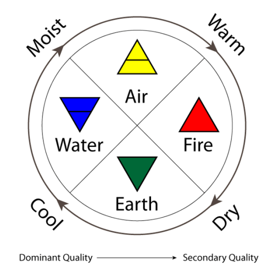
Interactive periodic table that clarifies the major divisions of the periodic table:
http://www.ptable.com/

Unit 1.3 - Atomic Theory (Hebden p 139 – 158)

1) Early Greek Theories of Matter

- As early as the $5th$ century B.C. Greek philosophers **Leucippos** and his student **Democritus** reasoned that if a sample of solid matter was repeatedly cut into smaller pieces, the eventual result would be a particle so small it couldn't be cut into anything smaller.
- This led to the belief that all substances were composed of small, invisible particles called "atoms" – the word *"atom" was derived from the* term "atomos" (Greek for uncuttable).
- At this time, atoms were believed (correctly) to be of different sizes, to have regular geometric shapes, and to be in constant motion.

- **Aritstole** proposed that all types of matter were made up of different proportions of four basic elements: earth, air, fire and water – this idea was accepted for almost 2000 years.
- Aristotle's model was questioned and deemed irrational with the scientific revolution of physics and the implementation of quantitative measurements.

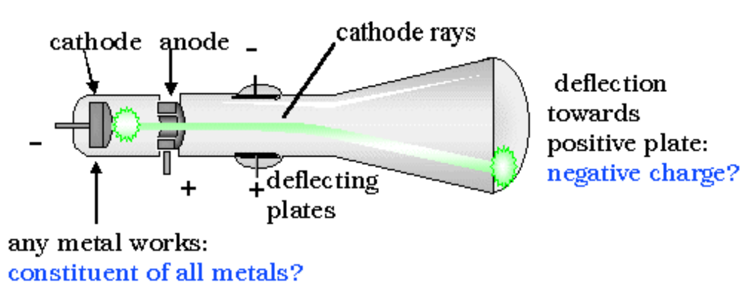


2) Dalton's Atomic Theory - 1808

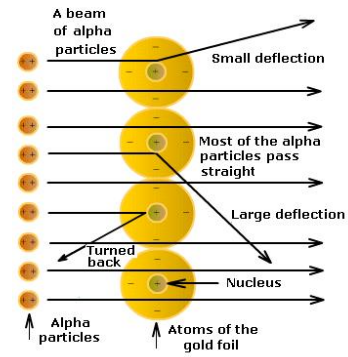
- All matter is composed of extremely small particles called atoms, which cannot be broken into smaller particles, created nor destroyed – *Law of conservation of mass.*
- The atoms of any given element all have identical properties and are different from the atoms of other elements.
- Atoms of different elements combine in specific ratios to form compounds.
- In a chemical reaction, atoms are separated, rearranged, and recombined to form new compounds.

3) Joseph John Thomson

- 1897 cathode-ray tube experiment – deflection of beam of particles toward the positive magnetic plate indicated that the particles had a negative charge – Thomson called these negative particles *ELECTRONS.*
- He also found that the electron was almost 1/2000 the size of a hydrogen ion - he had found a particle smaller than the smallest atom!



4) **Ernest Rutherford**



- 1909 student of J.J. Thomson's – created the gold foil experiment and deducted that an atom must contain a tiny positively charged core (1/10 000th of the total atom) and named it the *NUCLEUS.*
- Predicted the existence of the neutron – particle without an electrical charge (neutral particle) in the nucleus of an atom.
- Identified and named *PROTONS* as positively charged particles within the nucleus of an atom.

5) **H.G.J. Moseley**

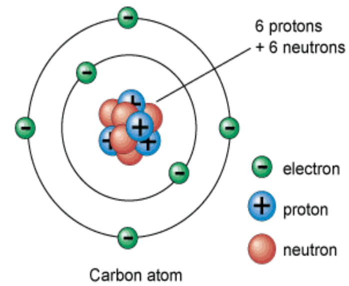
- His x-ray experiments showed that the positive charge in the nucleus of atoms increases by one unit in progressing from each element to the next – led to the discovery of atomic number.
- Because of this discovery, elements are now arranged in the periodic table in order of atomic number – i.e. by the number of protons they have in their nucleus (rather than by atomic mass).

6) **Neils Bohr**

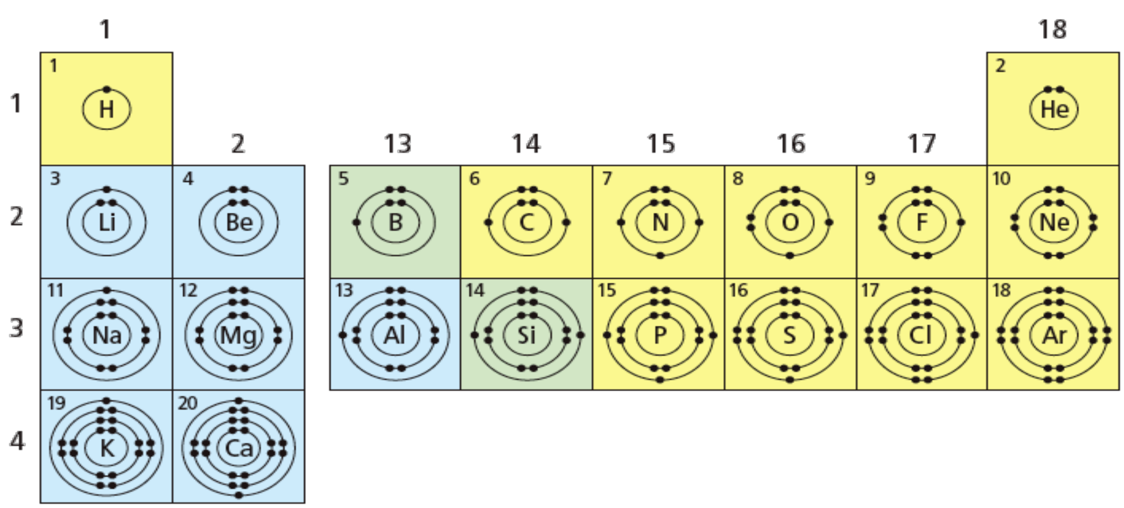
- 1913 - proposed that the electrons in an atom are restricted to specific energies and travel in paths called "orbits" at a fixed distance from the nucleus.
- Since then, his theory has been discarded but there are aspects of it that are useful for describing electron configuration:

i) The higher the energy level of an electron, the further it is from the nucleus.
ii) The maximum number of electrons in the first 3 energy levels is $2, 8\ and\ 8$.
iii) An atom with the maximum number of electrons in its outermost energy level is stable and unreactive.
iv) Bohr suggested that the properties of the elements can be explained by the arrangement of electrons in orbits around the nucleus.
e.g. All noble gases have extremely low chemical reactivity because they have full outer orbits.
e.g. All alkali metals are reactive, and tend to react the same way, because they have only 1 electron in their outer orbit.

Modern Day Atomic Theory



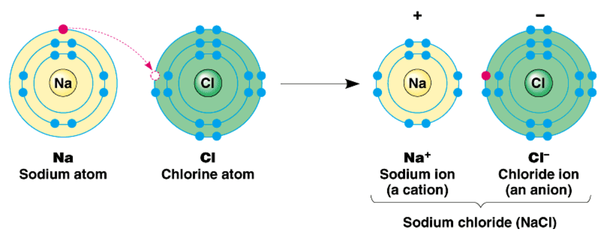
- An atom **is the smallest particle of an element** that still has the properties of that element.
- $50\ million$ atoms lined up end-to-end $= \ 1\ cm$
- Atoms are made up of smaller particles known as sub-atomic particles: *PROTONS (p, + charge), NEUTRONS (n, no charge) & ELECTRONS (e, - charge).*
- The *NUCLEUS* is at the *center of an atom* and is *composed of protons and neutrons.*
- The number of electrons around the nucleus of a *NEUTRAL atom* (no electrical charge), is equal to the number of protons in its nucleus.
    - Electrons exist in the space surrounding the nucleus in energy levels – each energy level can hold a specific number of electrons.
    - An *energy level* represents a specific value of energy of an electron. The number of occupied energy levels in any atom is "normally" the same as the period number in which the atom appears (true for all atoms in $periods\ 1-3$).
    - Energy levels must be filled with the maximum number of electrons it can hold before a new energy level is started.
    - For the first 3 energy levels, the maximum number of electrons that can be present are $2, 8, and\ 8$.
    - The electrons in the highest energy level of an atom (outermost energy level) are called *VALENCE ELECTRONS.*
    - *Stable atoms* have full valence shells.

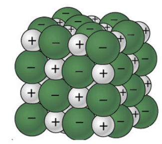


- Modern day chemists believe, that chemical reactions of elements and may compounds occur by the rearrangement of electrons.

Ions

- **Atoms that have either gained or lost electrons (have a + or – charge).**
- Metals tend to *LOSE* valence electrons, and end up as ions with a " + " charge (more protons than electrons).
- Non-metals tend to *GAIN* electrons and end up as ions with a "– " charge (more electrons than protons).





- Metal atoms tend to *LOSE* electrons and become ***CATIONS*** *- positively charged ions.*
- Non-metals tend to GAIN electrons and become **ANIONS** – *negatively charged ions.*
- *Once IONIC compounds (such as NaCl)* form, they are fairly *unreactive i.e. stable* because they have full outer energy levels.
- *General Rule – Atoms of the representative elements (groups 1, 2 and 13 to 18) form ions when losing or gaining electrons to form the SAME STABLE ELECTRONIC STRUCTURE AS ATOMS OF THE NEAREST NOBLE GAS.*

Size of Ions Relative to a Neutral Atom

1) **Negative Ions**

- The volume of negative ions increases as the number of repulsing electrons increases (like-charges repel), because of electrostatic repulsive forces between the electrons.

2) **Positive Ions**

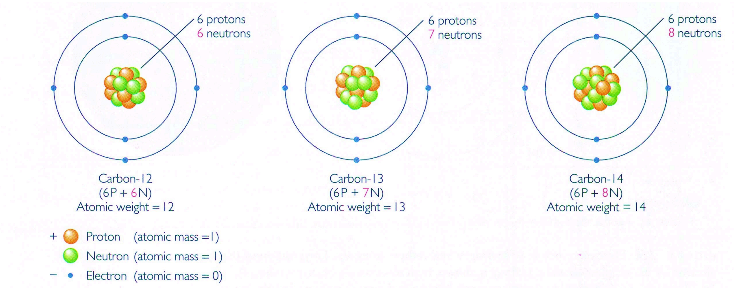
- The volume of positive ions decreases as the number of repulsing electrons decreases.
- As the number of electrons decreases, so does the amount of repulsion in the remaining electrons (each remaining electron has more space within the electron shells).

Special Cases of Ions

- Metalloids such as boron, carbon and silicon rarely form ions.
- Hydrogen atoms usually form positive ions by losing an electron – although it is unusual, negative hydrogen ions can be formed.
- Many of the transition metals do not always form ions with the same charge – many have multiple combining capacities – e.g. Iron may become a 3^+ cation or a 2^+ cation.

Isotopes

- Atoms of an element that have the *SAME atomic number* ($same\ \#\ of\ protons$), but a *DIFFERENT number of neutrons*.
- Different isotopes of the same element have the *same chemical properties*, but *different masses* (because each isotope has a different number of neutrons, and neutrons have a measurable mass).
- All elements exist naturally as a mixture of isotopes.



Atoms and Their Arrangement on the Periodic Table

- **Atomic Number –** the number of *PROTONS* in the nucleus of an atom.
  - Each element on the periodic table has a unique number of protons in the nucleus of its atoms.
  - The periodic table is organized according to the atomic number.

- **Mass Number**- the total number of protons plus neutrons in an atom - $\boldsymbol{mass\ \# = p + n}$.
  - Electrons are *not* involved when calculating the mass of an atom - their mass is so small it is said to be "negligible".

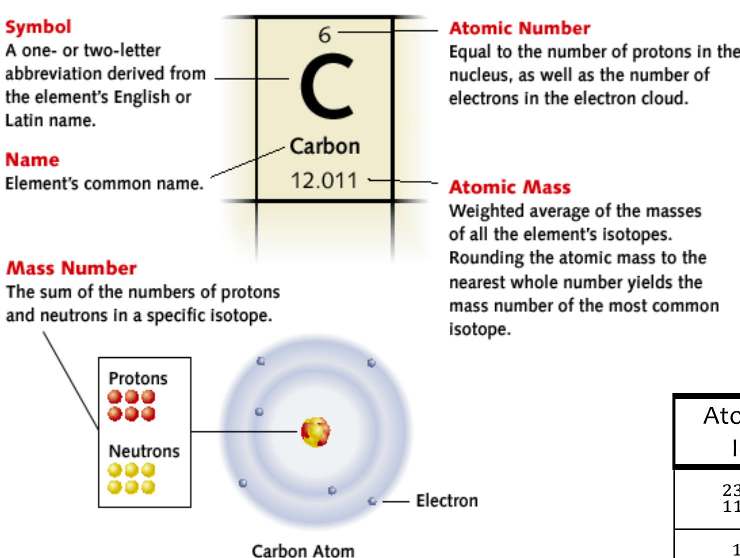
- Standard atomic notation for any given element follows these guidelines:

$$^{M}_{A}E$$

M = Atomic Mass (Neutrons + Protons)

A = Atomic Number (Protons)

E = Element



Atomic Mass is also commonly called:
"amu" – atomic mass units
"Molar Mass" – always written in grams per mol

Atom or Ion	Protons	Neutrons	Electrons
$^{23}_{11}Na$			
$^{1}_{1}H$			
$^{37}_{17}Cl$			
Al^{3+}			
Cl^{-}			

Periodicity and Atomic Theory

- By comparing the chemical reactivity of elements and compounds, patterns can be revealed within the periodic table for elements in different periods and families.

Ionization Energy (Forming Positive ions) & Electron Affinity (Forming Negative ions)

- The energy required to remove an electron from a neutral atom.
- As the force of attraction between the nucleus and electrons increases, more energy is needed to remove electrons from atoms.
- Fr has the lowest ionization energy and He has the largest ionization energy.

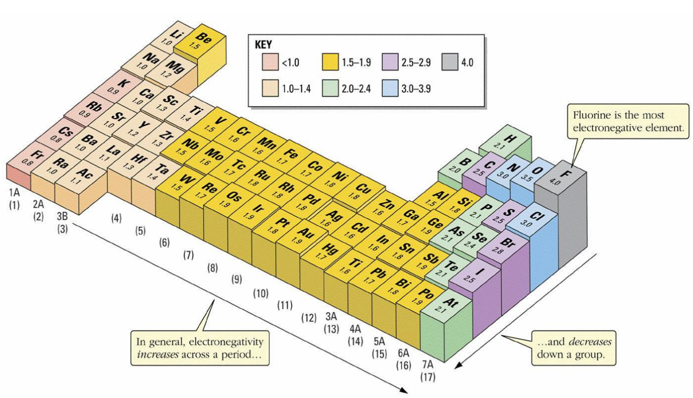
Atomic Radius

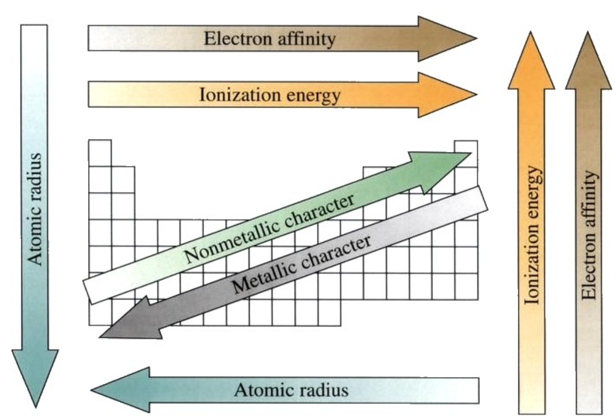
- Describes how tightly electrons are held to the nucleus.
- As you go down the periodic table, atomic radii of atoms *INCREASES*, this is a result of more electron shells each time you go down a period.
- As you go across the periodic table, atomic radii *DECREASES*, because there are more electrons in the shells and more protons in the nucleus causing a greater force of attraction that pulls the electron shells in closer to the nucleus.

Electronegativity

- The tendency atoms to attract electrons from neighbouring atoms.
- Atoms with high electronegativity have the strength to attract other electrons from neighbouring atoms.

  e.g. The halogens strongly attracted to its own valence electrons and form diatomic molecules - F_2, Cl_2, Br_2 etc.
- Fluorine has the largest electronegativity and Francium has the lowest.





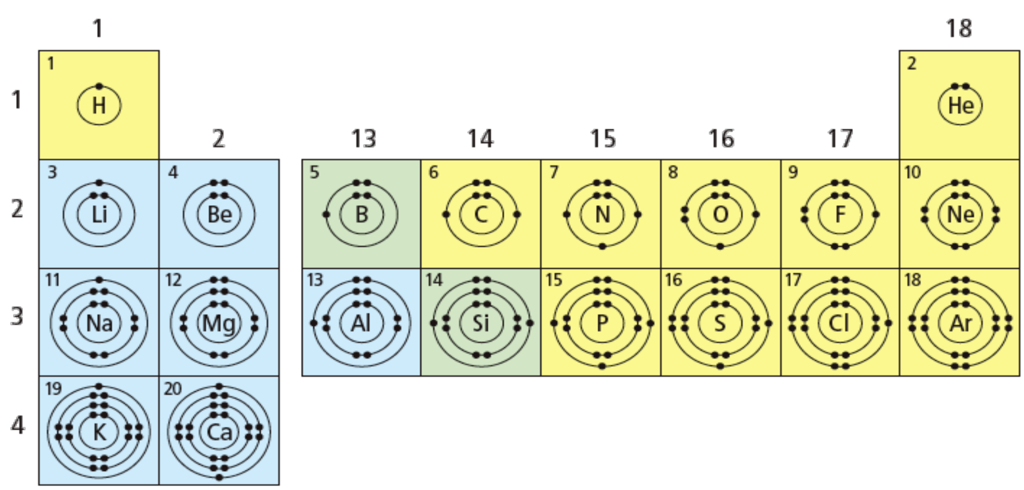
Remember the following when examining trends down or across the periodic table:

1) When going *DOWN* a group, properties of elements are affected by the *increasing size* of the atoms and the *increasing distance* between the nuclei of the atoms and their valence electrons.

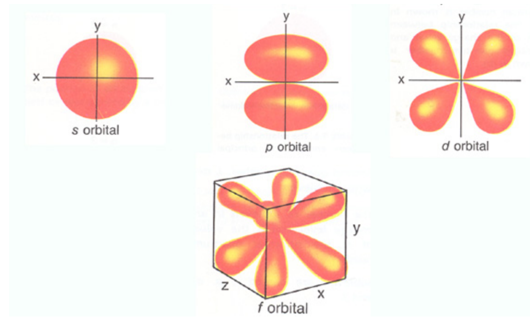
2) When going *ACROSS* a period, properties of elements are affected by the *differing number of valence electrons, nuclear charge (# of protons) and charge on the atom.*

Bohr Diagrams & Electron Configuration

- Bohr diagrams show how many electrons appear in each electron shell around an atom for the first three periods.
- All elements in a period have the same number of electron shells.
- Each shell holds a maximum number of electrons.
- Electrons in the outermost shell are called *valence electrons.*
- The period # = # of shells in an atom.
  e.g. **Argon** atomic number = 18 – Argon needs 3 electron shells to hold 18 electrons, so it is in $period\ 3$.
- The last digit of the group # = # of electrons in the valence shell (True for the representative elements, $groups\ 1, 2\ and\ 13\ to\ 18$ – does not hold true for the transition metals in $groups\ 3\ to\ 12$).



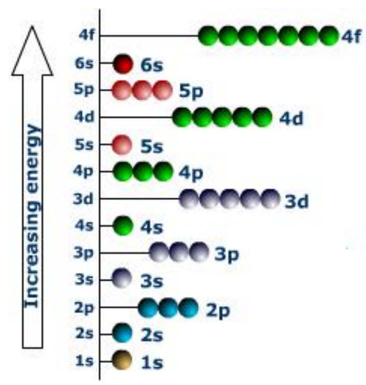
- In reality, within each shell electrons are in constant motion and move in wave-like patterns that have the shape of 3D clouds– they *DO NOT* move around 2D rings as commonly depicted in Bohr diagrams.

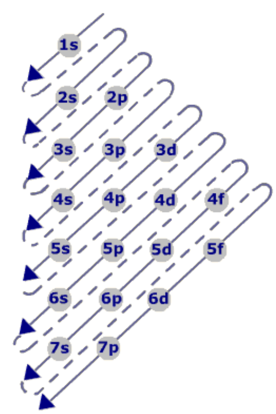


- These 3D "clouds" are called orbitals:
  - each s orbital holds a maximum of 2 electrons
  - each p orbital holds a maximum of 6 electrons
  - each d orbital holds a maximum of 10 electrons
  - each f orbital holds a maximum of 14 electrons

- According to the Theory of Quantum Mechanics, an orbital is defined as a region of space in which there is a 95% probability of finding an electron with a given energy.
- Electrons appear in shells in a very predictable manner – they fill the lowest energy levels within each orbital first before they start filling energy levels in the next orbital.
  - There is a maximum of 2 electrons in the first shell → $1s^2$
  - 8 in the 2nd shell → $1s^2 2s^2 2p^6$
  - 8 in the 3rd shell → $1s^2 2s^2 2p^6 3s^2 3p^6$

e.g. The electron configuration of Na is:





Unit 1.4 – Chemical Bonding & Forming Compounds

- When two atoms get close together, their valence electrons interact.
- If the valence electrons can combine to form a low-energy bond, a compound is formed.
- Unstable atoms react to form bonds with other atoms so they have the same number of valence electrons as the nearest noble gas (recall, noble gases are stable unreactive elements).
- All chemical bonding is based on the electrostatic relationships between atoms in a molecule:

  i) *Opposite charges attract* each other.

  ii) *Like Charges repel* each other.

  iii) The *greater the distance* between the 2 charged particles, the smaller the attractive or repulsive force between them.

  iv) The *greater the charge* on 2 particles, the greater the force of attraction or repulsion between them.

Types of Chemical Bonds

1) **Ionic bonds**

- Form when an *electron(s) are transferred from a metal to a non-metal.*
- The electron transfer creates a positive cation (metal) and a negative anion (non-metal) atom.
- The electrostatic attraction between the " +" and "– " ions is what creates an *IONIC BOND* between metal and non-metal ions.
- Ionic bonds are *very strong*, consequently compounds held together by ionic bonds have high melting points. i.e. lots of energy is required to break ionic bonds.
- Ionic bonds formed between metals and non-metals and create repeating molecular patterns – crystals.

E.g. lithium and oxygen form an ionic bond in the compound Li_2O

- Helpful patterns to recognize regarding the formation of ions:

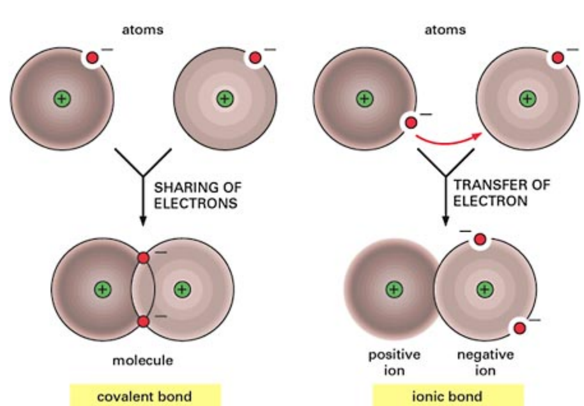
$Group$ 1 elements form $+1$ ions	$Group$ 16 elements form -2 ions
$Group$ 2 elements form $+2$ ions	$Group$ 17 elements form -1 ions
$Group$ 15 elements form -3 ions	Noble gases are stable and do not form ions

2) **Covalent bonds**

- Form when *two or more non-metals share a pair of valence electrons.*
- Sharing electrons allows both non-metal atoms to have full valence orbitals - stable molecule.
- Covalent bonds form when both atoms involved have high electronegativities.
- Covalent bonds form when non-metal atoms attract each other's electrons strongly but will not let go of their own electrons.
- Electrons stay with their atom, but overlap with other shells.

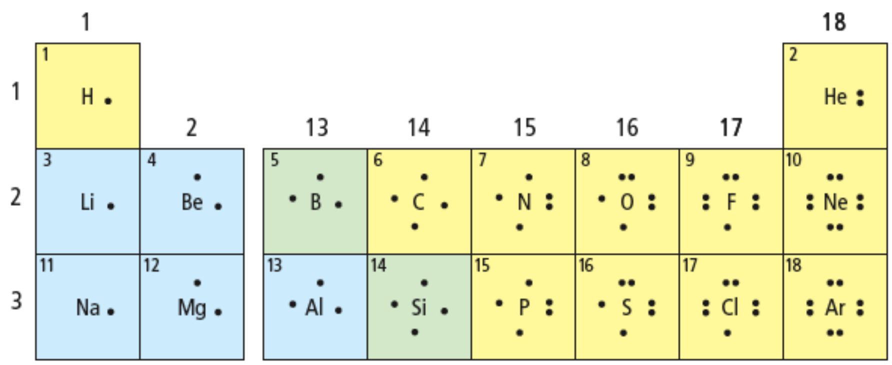
E.g. carbon and oxygen form the covalent compound CO_2 - each oxygen atom shares one of its electrons with carbon to create two covalent bonds that hold a CO_2 molecule together.

- single covalent bond - 1 shared pair of electrons
- double covalent bond - 2 shared pairs electrons
- triple covalent bond - 3 shared pairs electrons
- the bond length of a triple bond < double bond < single bond



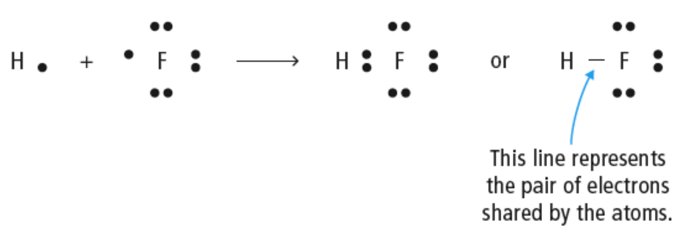
Lewis Diagrams

- Lewis diagrams model the arrangement of *valence electrons* in atoms, and explain and predict empirical formulas of compounds (whereas Bohr diagrams model the full electron configuration).
- Dots representing electrons are placed around the element symbols at the points of the compass (north, east, south, and west).
- Electron dots are placed singly, until the fifth electron is reached, then they are paired until all valence electrons are accounted for.



Lewis Diagrams of Covalent Bonds

- All atoms (except hydrogen) wish to have a full valence shell or achieve an octet (8 electrons) - hydrogen only *requires* 2 *electrons* to have a full valence shell.
- The shared pairs of electrons are usually drawn as a straight line.



Steps to Drawing Lewis Diagrams:

1) Find out the total number of valence electrons you ***HAVE*** from *ALL* atoms in the molecule.

2) Figure out how many octet electrons the molecule ***NEEDS***, using the octet rule:
   - All atoms want eight valence electrons, so they can be like the nearest noble gas.
   - *Exceptions to the octet rule:*
     i) Boron - wants six valence electrons
     ii) Hydrogen – wants two valence electrons

3) Determine how many ***SHARED PAIRS (BONDS)*** are in the molecule.
   - Subtract the valence electrons from octet electrons: i.e. subtract the number you found in #1 above from the number you found in #2 above.
   - Divide this number by two (because each bond represents a 2 electrons).
   - The answer you get is the number of bonds in the molecule.

4) Draw an arrangement of the atoms for the molecule that contains the number of bonds you found in #3 above: Some handy rules to remember are these:

i. The most electronegative atom is always at the center of the molecule.
ii. Hydrogen and the halogens bond once.
iii. The group oxygen is in ($group$ 16) bonds twice.
iv. The group nitrogen is in ($group$ 15) bonds three times. So does boron.
v. The group carbon is in ($group$ 14) bonds four times.

- **NOTE:** you may have to assign double or triple bonds to achieve the octet with the available number of valence electrons.

5) Find the number of lone pair (nonbonding) electrons by subtracting the bonding electrons (#3 above) from the valence electrons (#1 above).

- Arrange these around the atoms until all of them satisfy the octet rule.
- Remember, ALL elements EXCEPT hydrogen want eight electrons around them, total. Hydrogen only wants two electrons.

6) Tidy up: replace each pair of electrons engaged in a bond with a dash, "—"

e.g. H_2O NH_3 F_2

Lewis Diagrams for Ions:

- For positive ions, one electron dot is removed from the valence shell for each positive charge of the ion.
- For negative ions, one electron dot is added to each valence shell for each negative charge of the ion.
- Square brackets are placed around each ion to indicate transfer of electrons.

e.g. NH_4^+ CO_3^{2-}