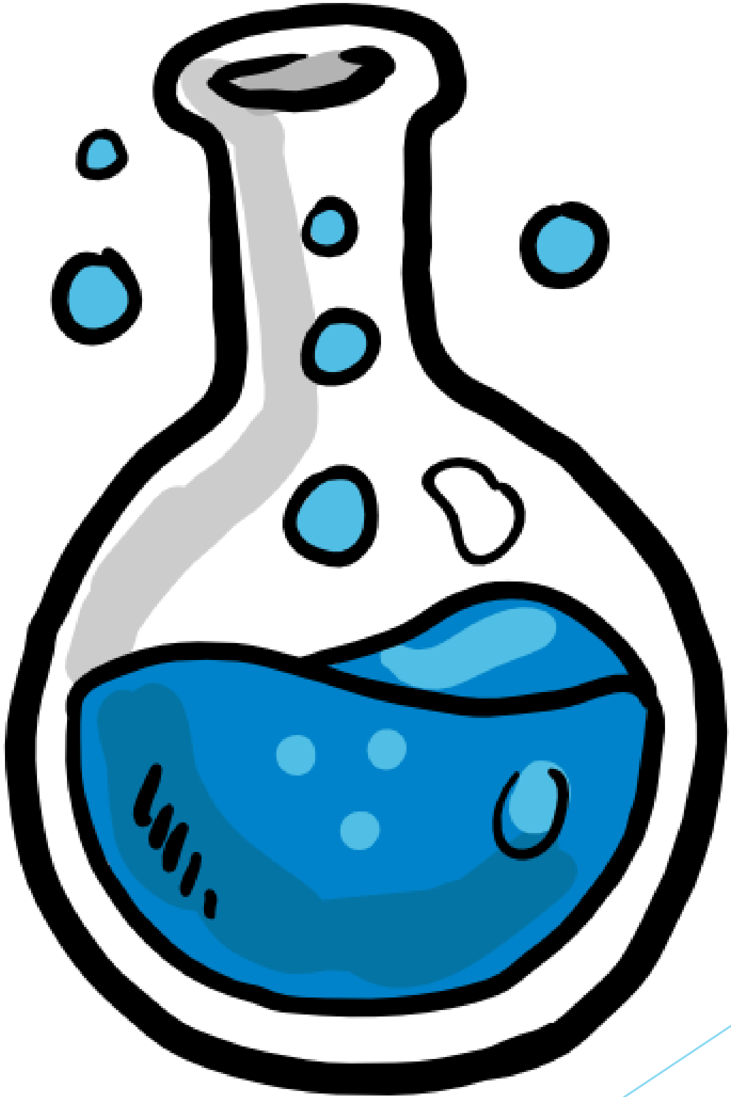
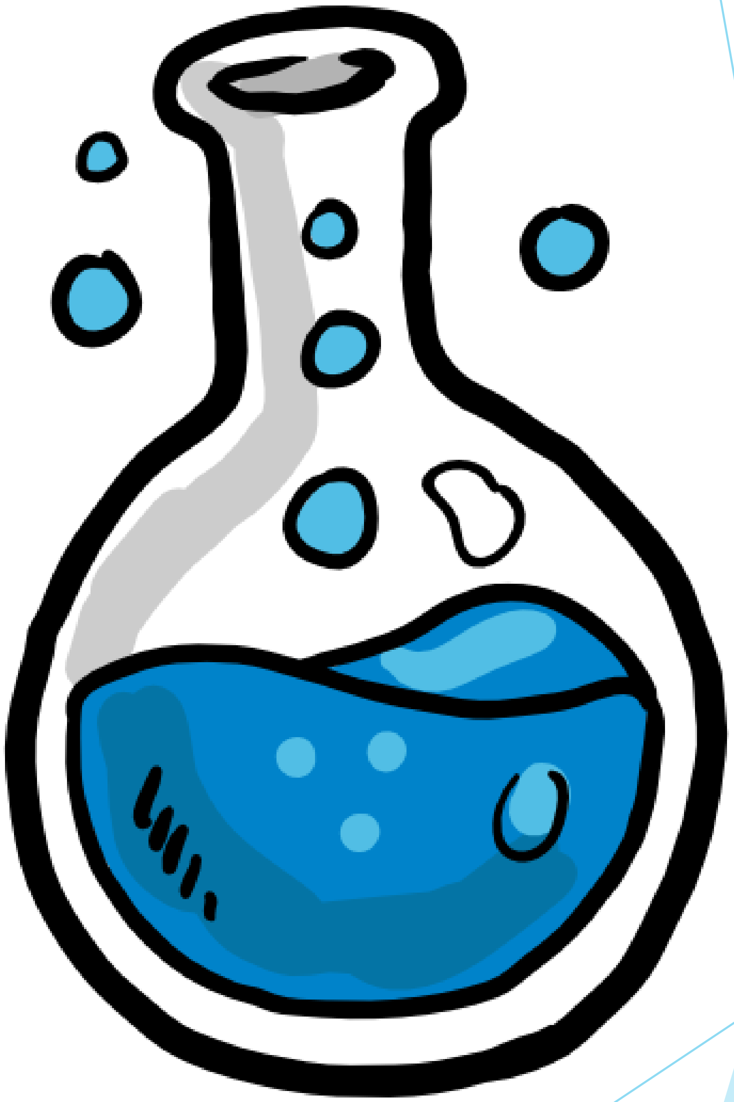
11 CHEMISTRY Module 1: Revision



Properties of Matter



Physical Properties

Can include:

- Homogenous (the same) or heterogeneous (different) compositions
- Colour
- Magnetism (alloys containing iron, nickel or cobalt)
- Particle Size
- Melting Points (lowest temp, solid → liquid)
- Boiling Points (volatile means easily converted to vapour)

Density

- Defined as the mass per unit of volume
- Typically expressed in Grams per Millilitre (g/mL) or kilograms per cubic metre (kg/m^3)
- Calculated by:

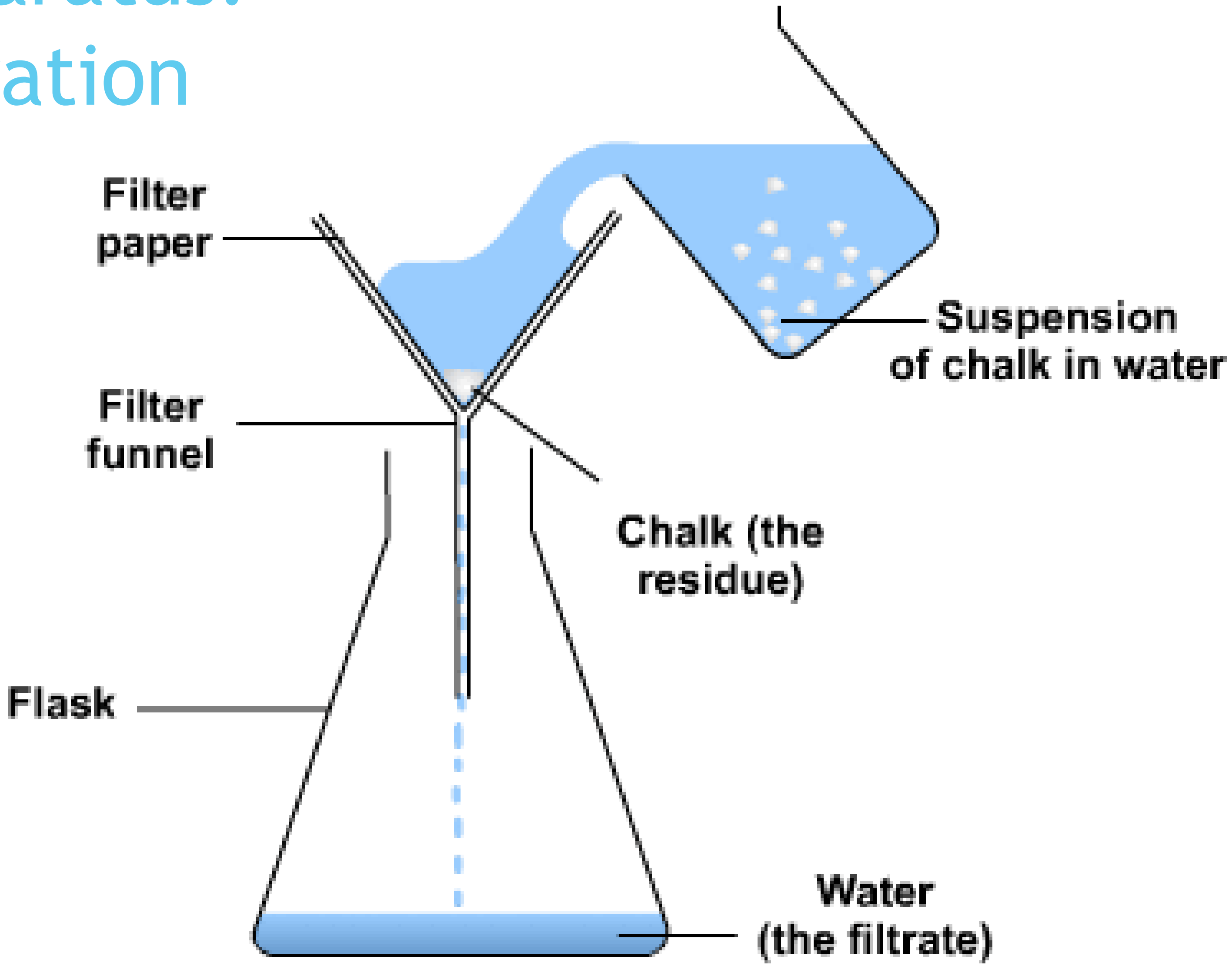
$$density = \frac{mass}{volume}$$

Methods of Separation:

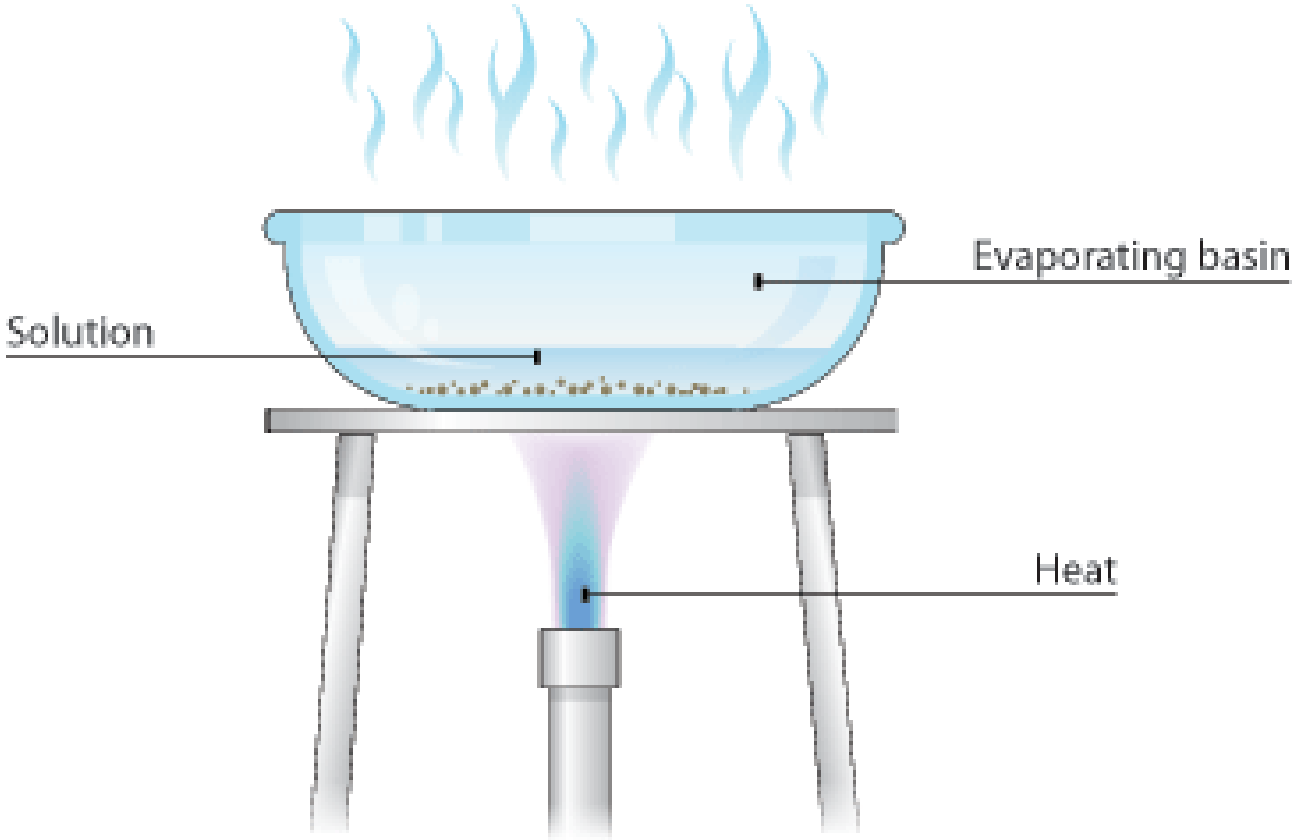
Separation of a mixture can be performed by:

- Filtration– can be used to separate insoluble solids from liquids, based on particle size. Can be used to collect insoluble solid or filtrate
- Based on differing solubility– two solids can be separated if one is soluble in a solvent, and one is not
- Boiling Points –
  - Evaporation (to dryness) – separating a dissolved solid (solute) from a solution
  - Distillation – Boiling a mixture of liquids, where the solute is non-volatile, and the volatile vapour is condensed and collected as distillate
  - Fractional Distillation – Separation of two volatile liquids based on differing boiling points
- Density – Allowing sedimentation of heavier particles and collecting off the liquid through decantation (pouring off top liquid layer)

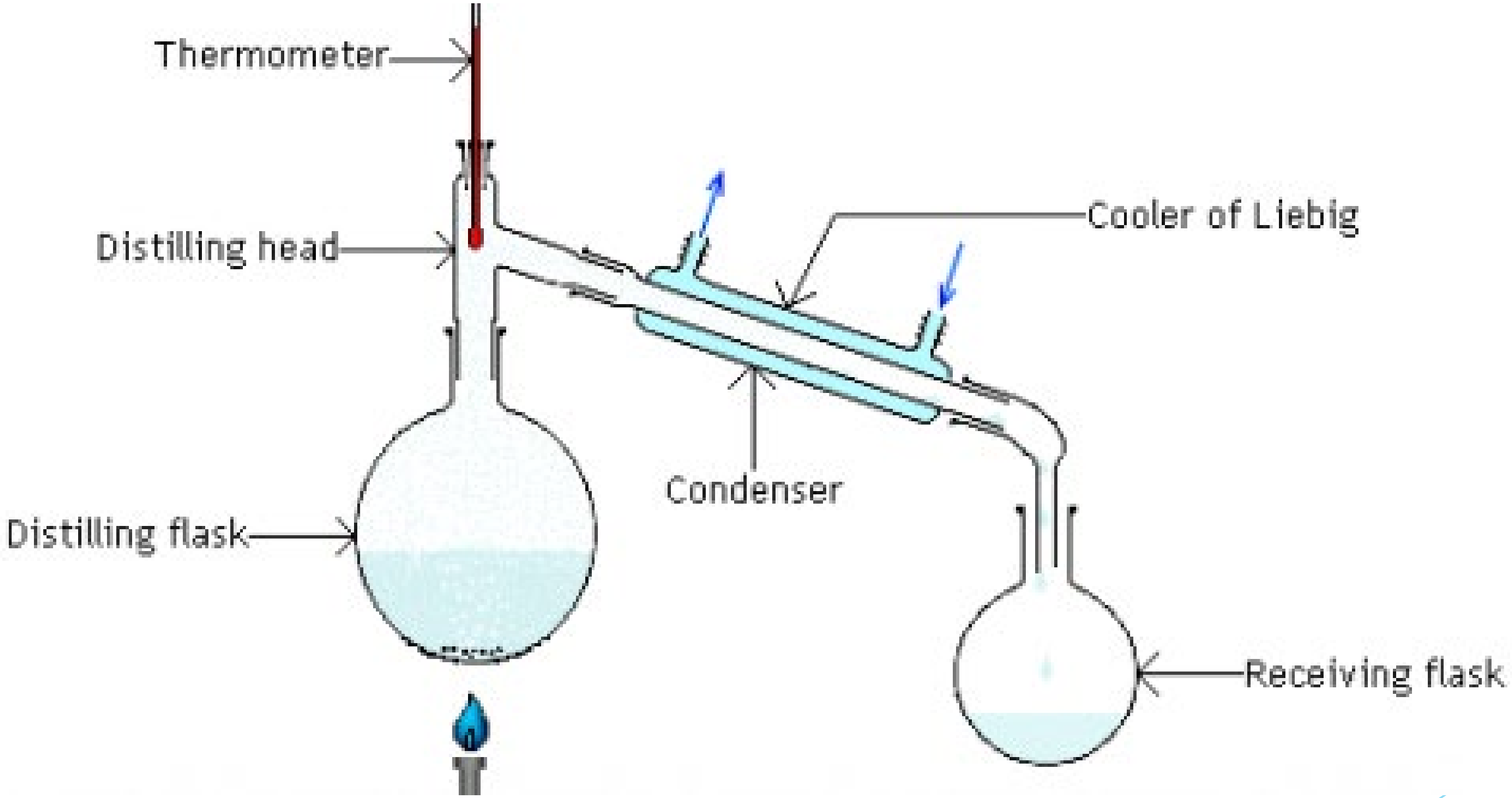
Apparatus: Filtration



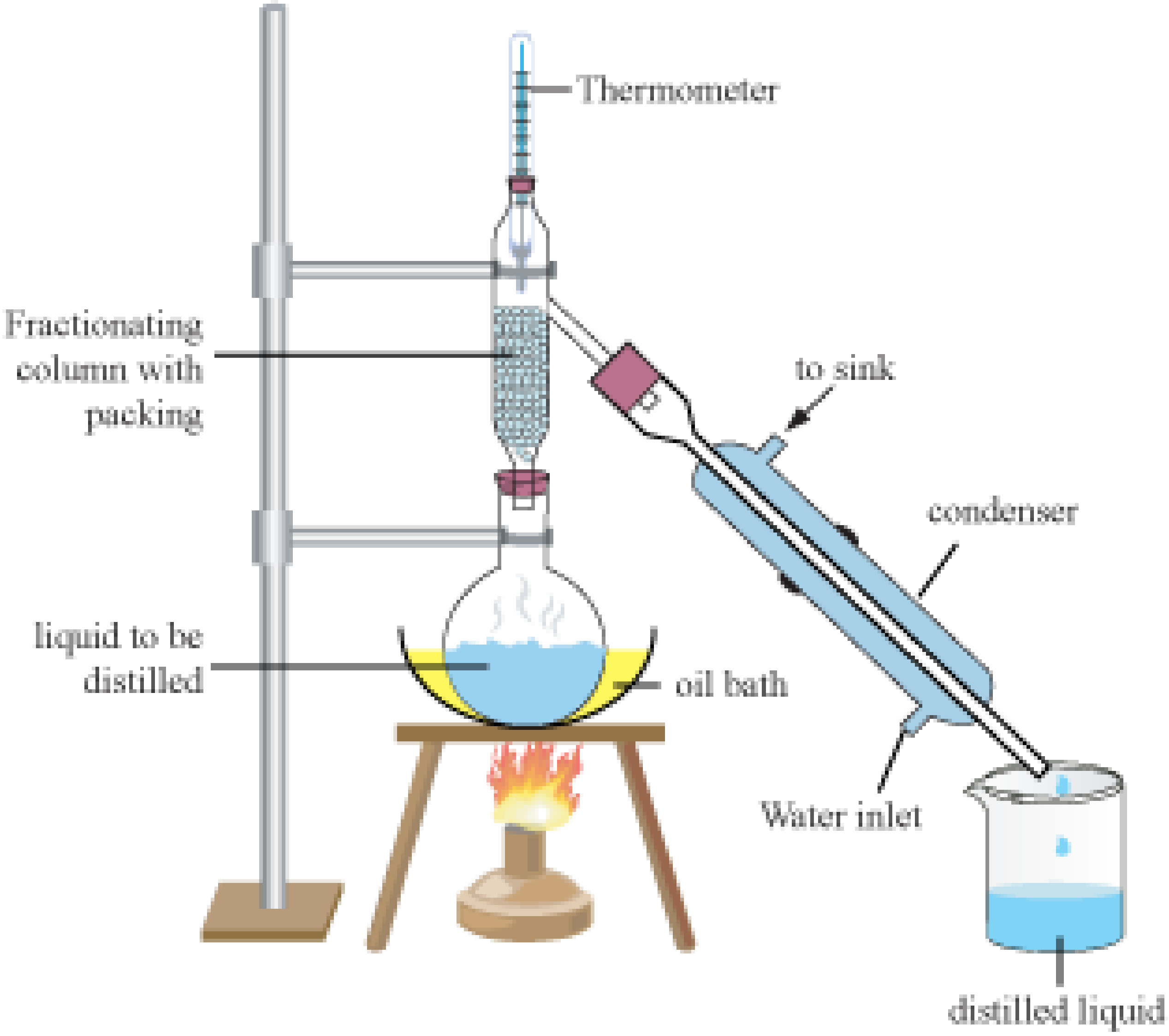
Apparatus: Evaporation



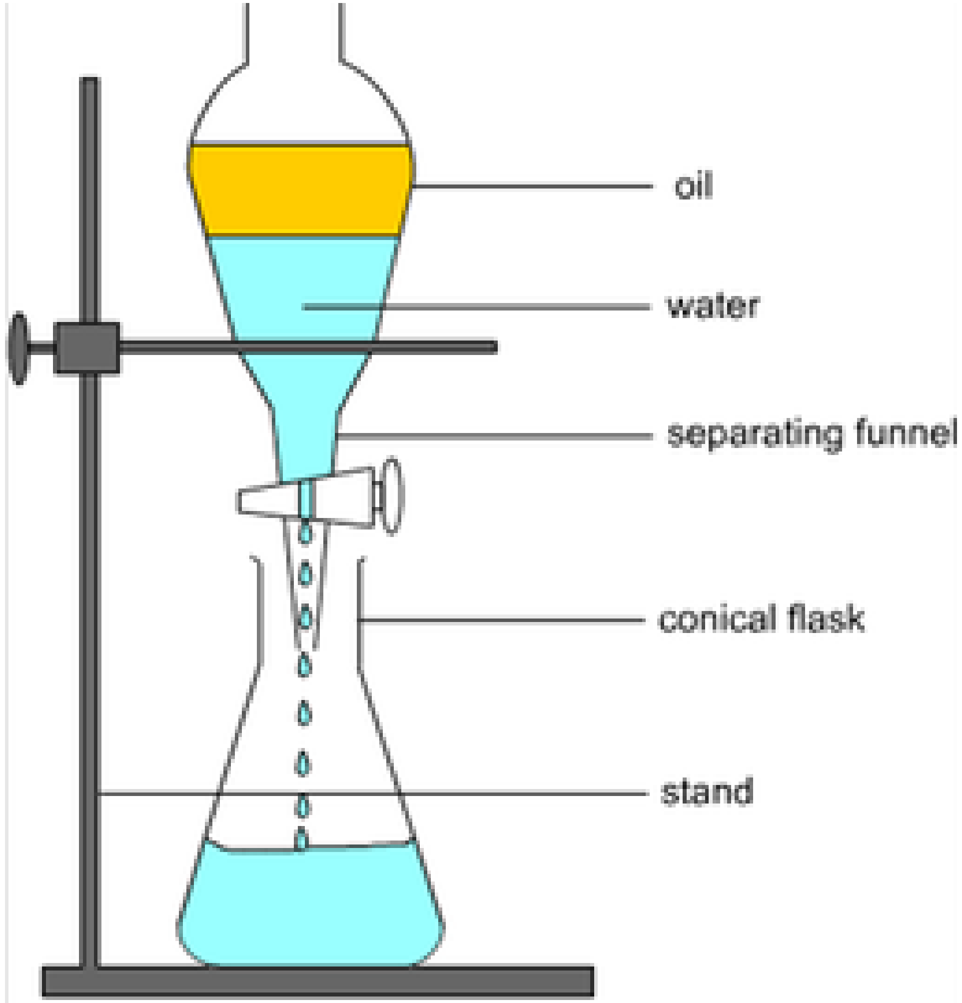
Apparatus: Distillation



Apparatus:
Fractional Distillation



Apparatus:
Separating Funnel



Gravimetric Analysis

- Analysis of a mixture be separating the substances and weighing them
- Often expressed as a percentage

$$Percentage\ Composition\ \% = \frac{Experimental\ sample}{Theoretical\ Sample}\ x\ 100$$

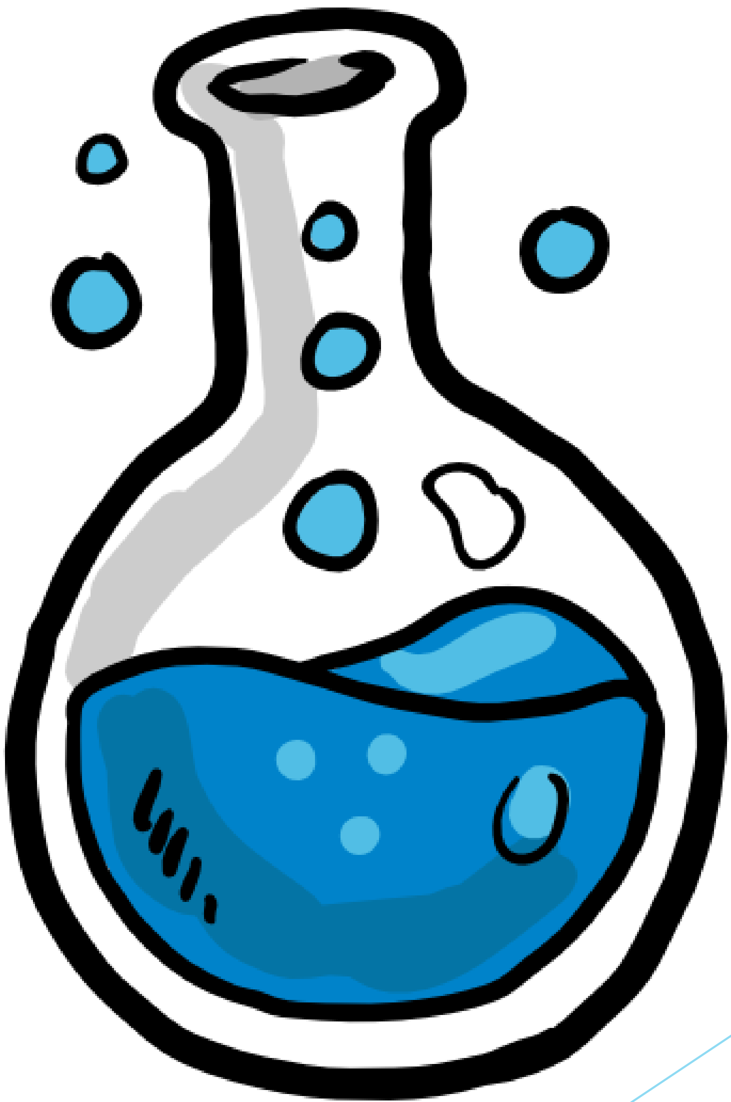
The Periodic Table

	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	
1 H 1.008																	2 He 4.003
3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 23.00	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.90	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.70	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
55 Cs 132.9	56 Ba 137.3	57 La 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.9	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.0	89 Ac 227.0	104 Rf (261)	105 Ha (262)	106 Unh (263)	107 Uns (262)		109 Une (267)									

Lanthanides	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (145)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
Actinides	90 Th 232.0	91 Pa 231.0	92 U 238.0	93 Np 237.0	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

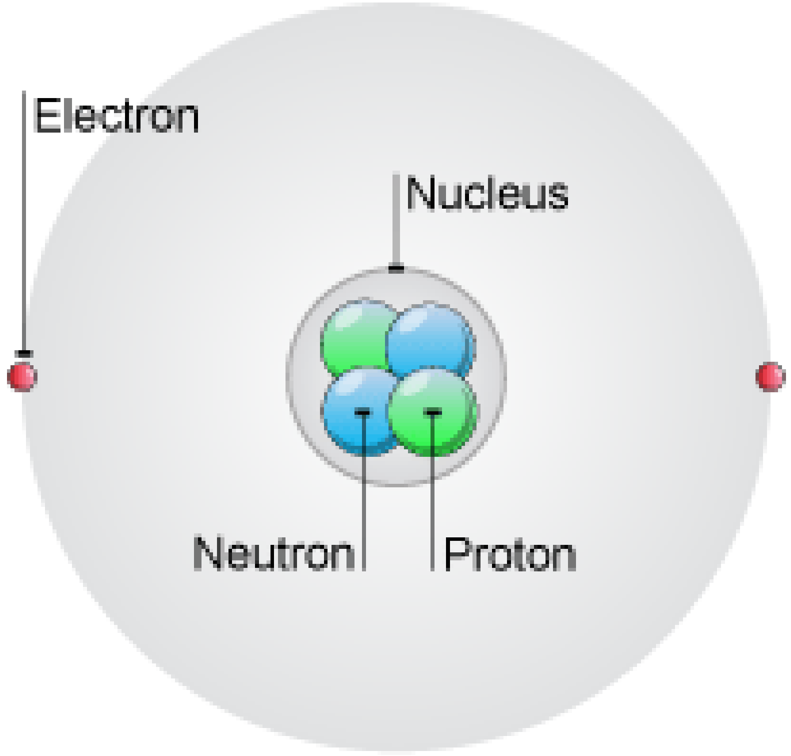
- Metals are elements that:
  - Solids at room temperature
  - Shiny appearance
  - Good conductivity
  - Malleable
- Most other elements are non-metals
- Columns = groups – same number of valence electrons in outside shells
- Rows = periods – same number of shells

Atomic Structure



Structure of an atom

- Subatomic Particles:
  - Electrons (e) = -1 charge
  - Protons (+ or p) = +1 charge
  - Neutrons (n) = 0 charge
- Nucleus consists of protons (atomic number) and neutrons (atomic mass- atomic number)
- Shells outside the atom hold electrons around the outside of the nucleus

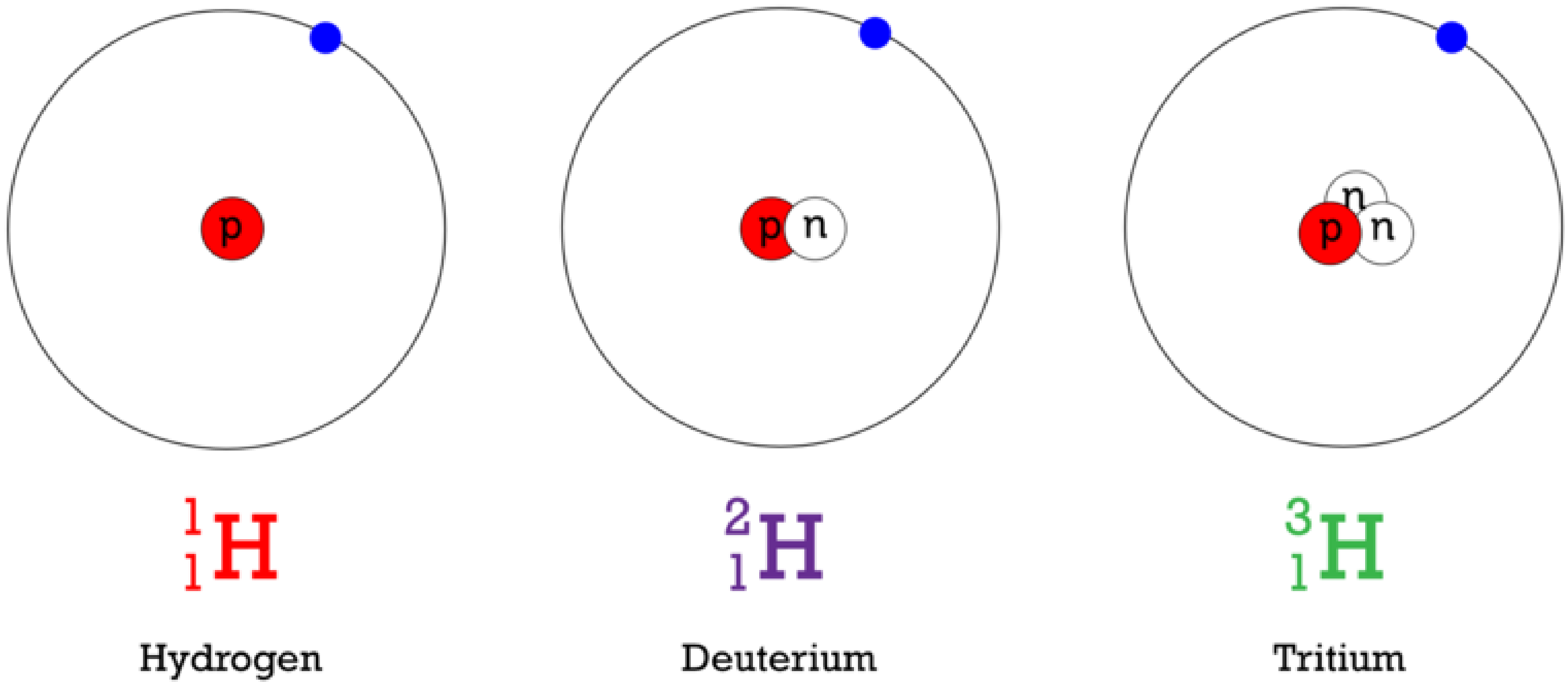


Isotopes

- Defined as ‘atoms of the same element that have different numbers of neutrons in their nuclei’
- ‘Relative abundance’ of an isotope is the percentage of that isotope in naturally occurring elements
- Isotopes of the one element have the same chemical properties and very similar physical ones.
- Symbol follows format:

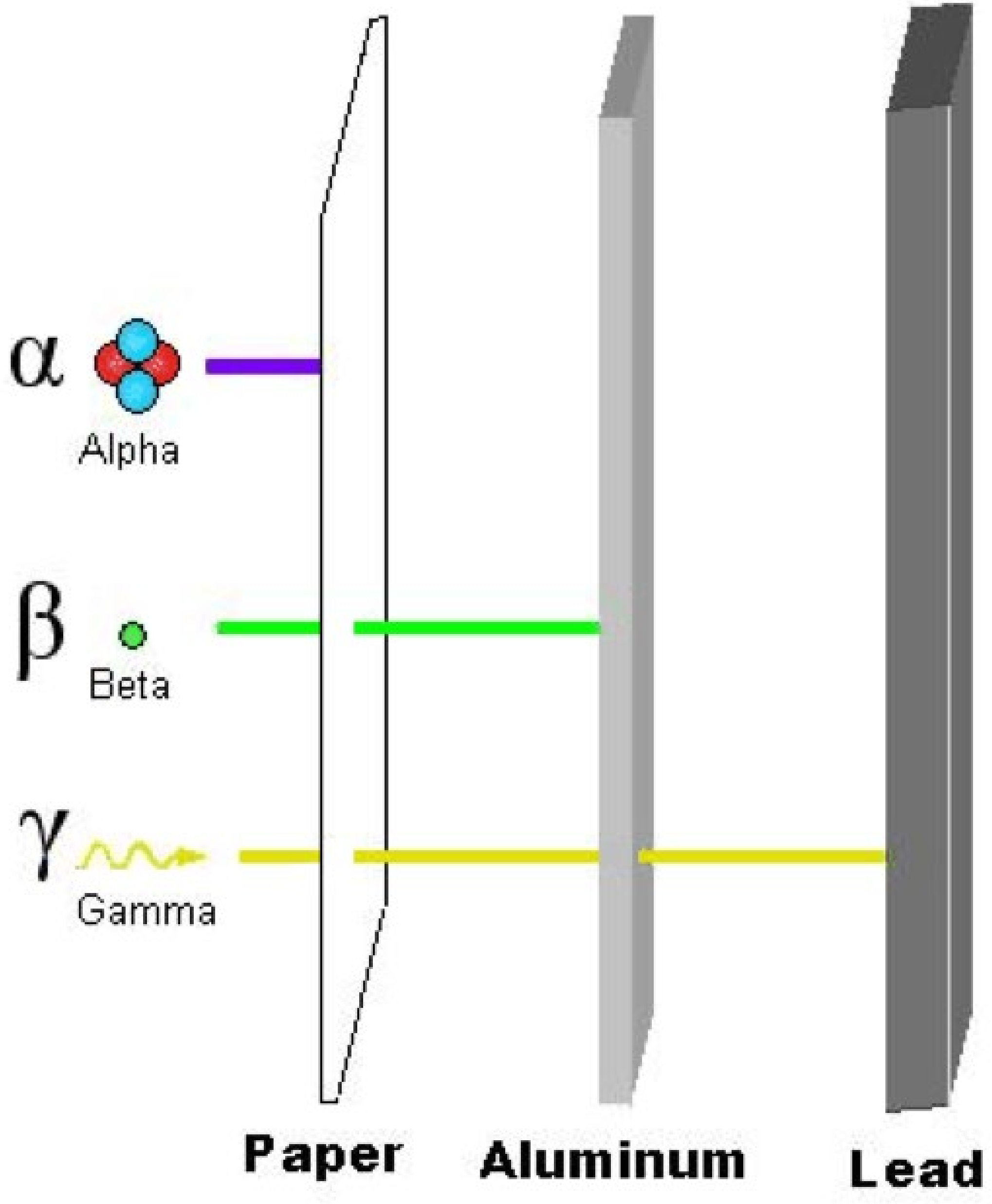
$$\frac{A}{Z}M$$

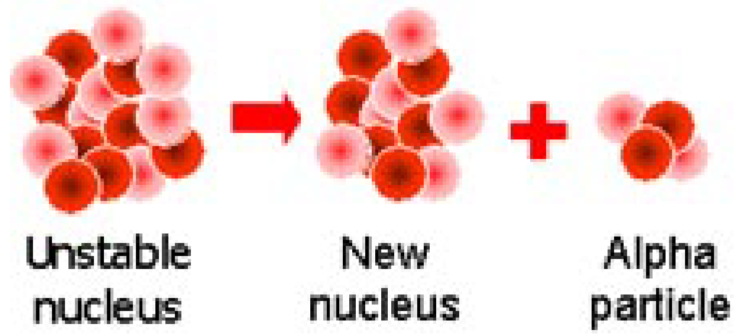
where A = Atomic mass (protons + Neutrons), Z = Atomic Number (protons) and M is the elemental symbol


p
p n
n
p n
^1_1H
^2_1H
^3_1H
Hydrogen
Deuterium
Tritium

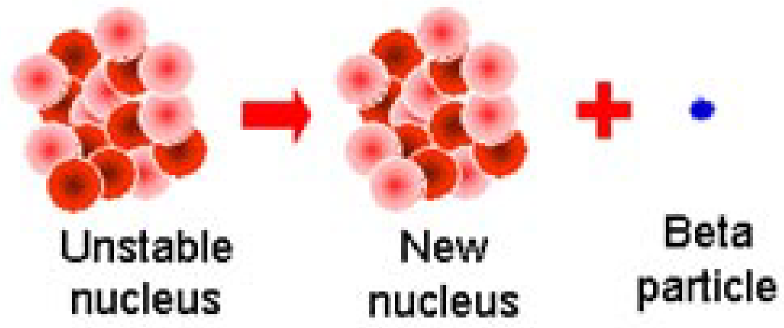
Unstable Isotopes

- Radioactivity is the spontaneous emission of radiation that occurs with certain isotopes and arises because some isotopes are unstable
- Three types of radiation are:
  - Alpha (helium particle)
  - Beta (electron)
  - Gamma (similar properties to X-rays)

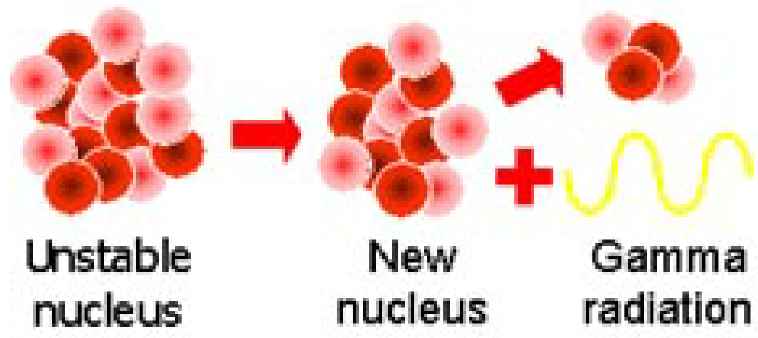




Alpha (α): atom decays into a new atom & emits an alpha particle (2 protons and 2 neutrons: the nucleus of a helium atom)



Beta (β): atom decays into a new atom by changing a neutron into a proton & electron. The fast moving, high energy electron is called a beta particle



Gamma (γ): after α or β decay, surplus energy is sometimes emitted. This is called gamma radiation & has a very high frequency with short wavelength. The atom is not changed

Electron Configuration

- Electrons live in energy levels, which is a cloud of probability for the location of the electron
- The arrangement of the electrons into shells or energy levels is referred to as the electron configuration
- Rules of 2.8.8. applies, before putting 2 electrons into the fourth shell and then fill the third shell up to 18.
- This makes Titanium (22 electrons) have an electron configuration of 2.8.9.2, and explains the arrangement and activities of the transition metals
- The driving force behind chemical reactivity is the force of the atom seeking to achieve 'stable electron configurations' ie a full outside shell
- The electrons in the outside shell (highest energy level) are called valence electrons.

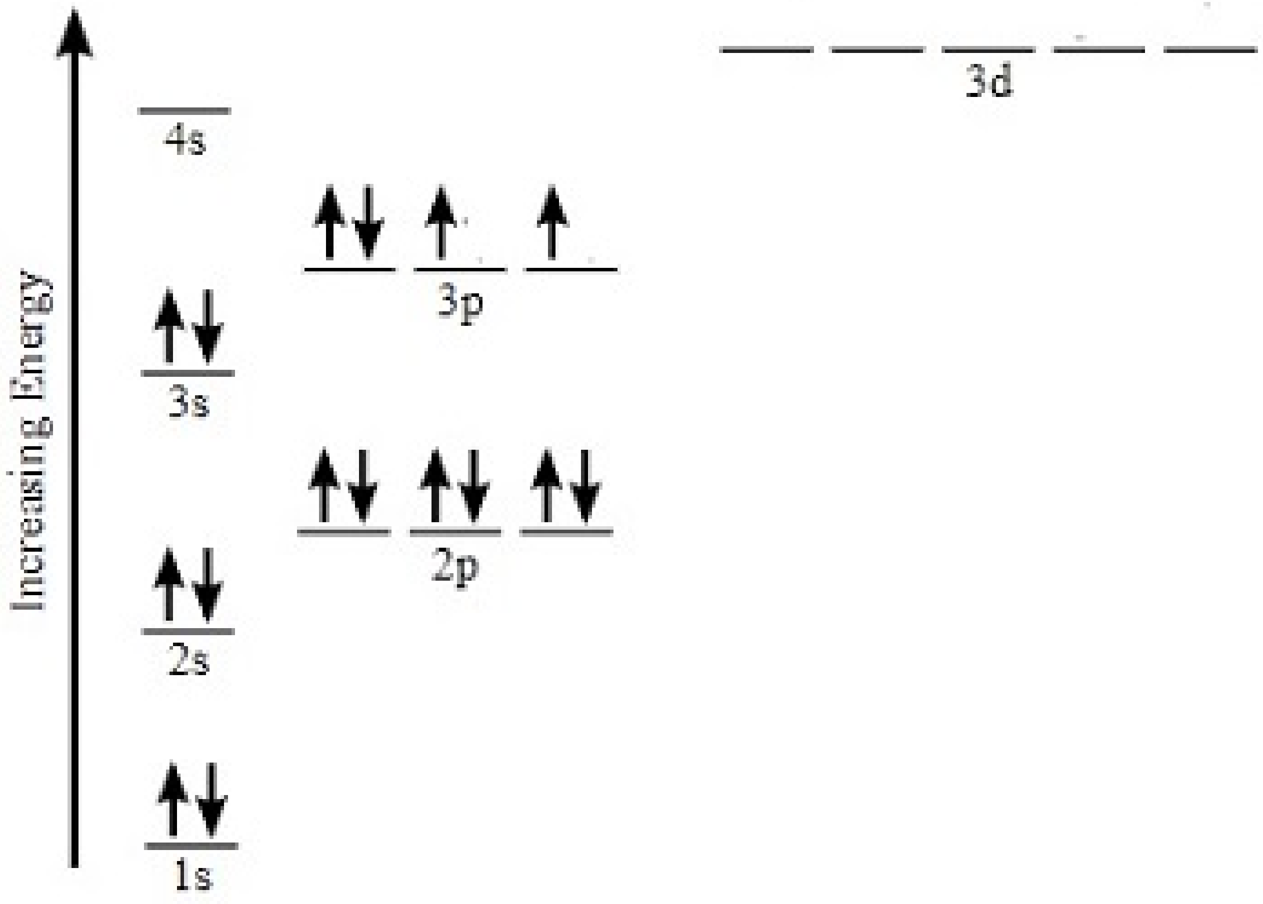
Orbitals

- Defined as ‘a volume of space surrounding the nucleus of an atom through which one or two electrons may randomly move’
- Each main energy level is made up of sublevels, called the s, p, d and f sublevels
- Each orbital can contain a maximum of 2 electrons
  - *s* can hold 2 electrons
  - *p* can hold 3 sets of 2 electrons
  - *d* can hold 5 sets of 2 electrons
  - *f* can hold 7 sets of 2 electrons
- Filling Order: We fill 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d

Example: Nitrogen

- Electron configuration 2.5.
- Orbital Notation: $1s^2\ 2s^2\ 2p^3$

Orbital Notation Diagram (arrows in boxes)



Atomic Emission Spectroscopy (AES)

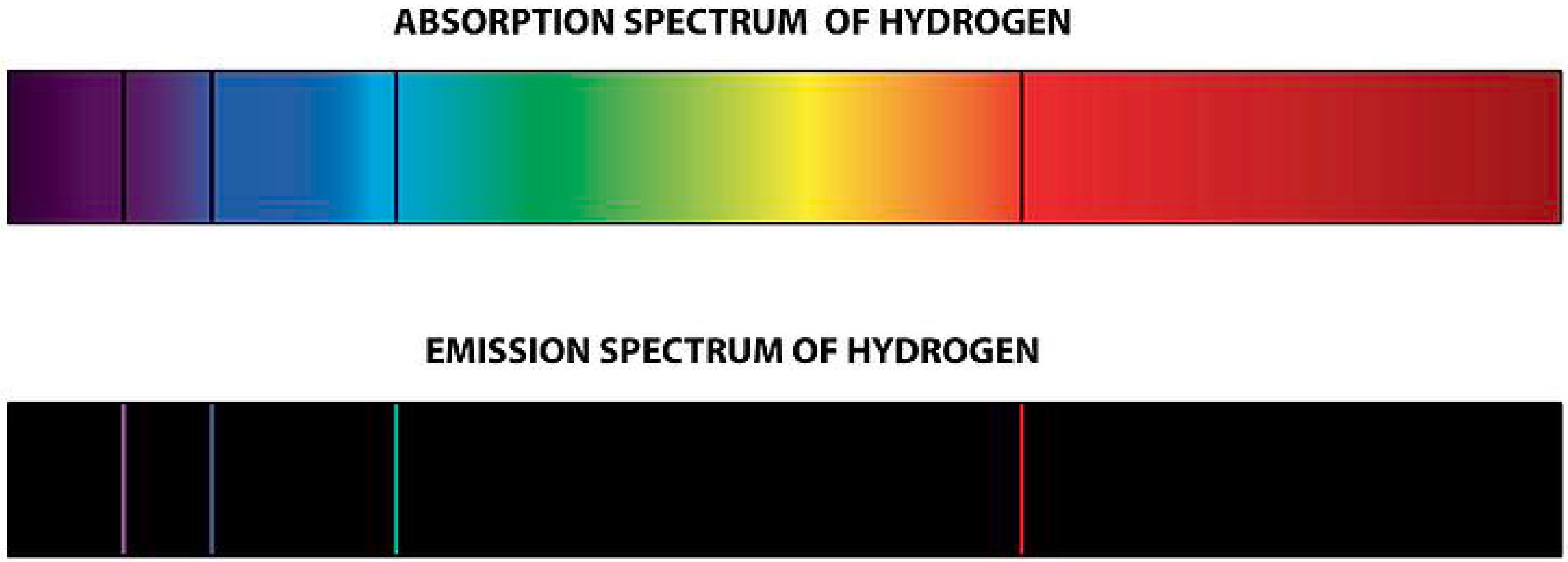
- Giving atoms extra energy (eg heating >1500°C) will cause electrons to jump up energy levels
- After a short time, the electrons fall back to their correct configuration, releasing the energy as emitted light – visible, UV or IR.
- The energy released is the same as the amount of energy absorbed to reach the 'excited state'
- The relationship between energy released and the wavelength is inversely proportional, related by Planck's Constant

$$\Delta E = \frac{hc}{\lambda}$$

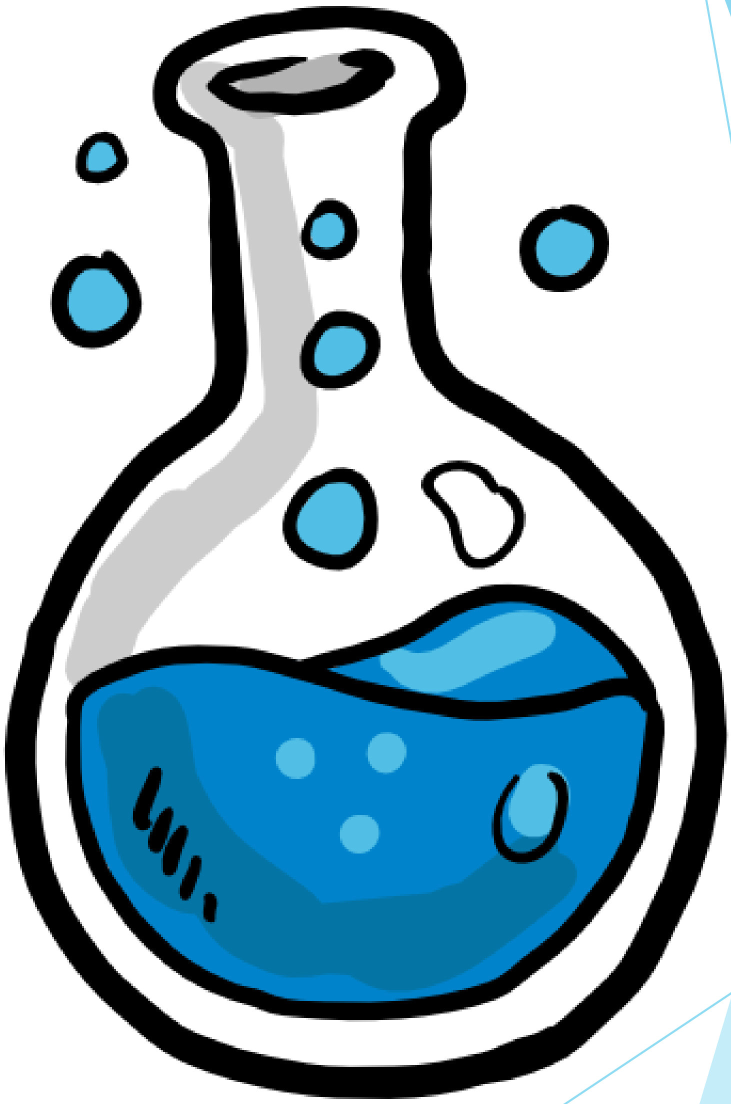
where *h* is Planck's constant and *c* is the velocity of light

Emission Specta

- Specific to each element, similar to a fingerprint
- Wavelength released corresponds to the energy required to excite a particular electron in an atom to the excited state, and this is equal to the wavelength within the visible spectrum



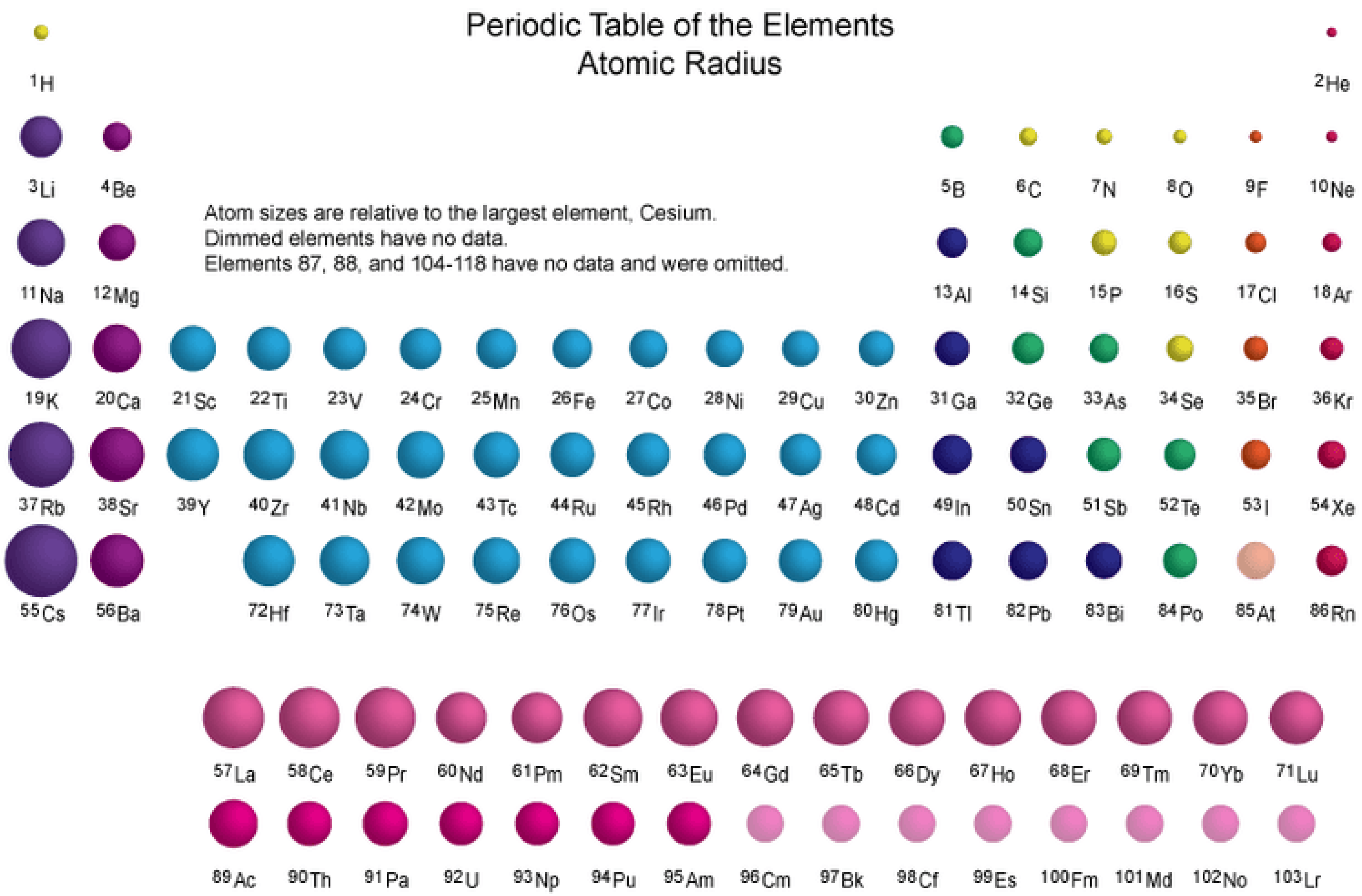
Periodicity



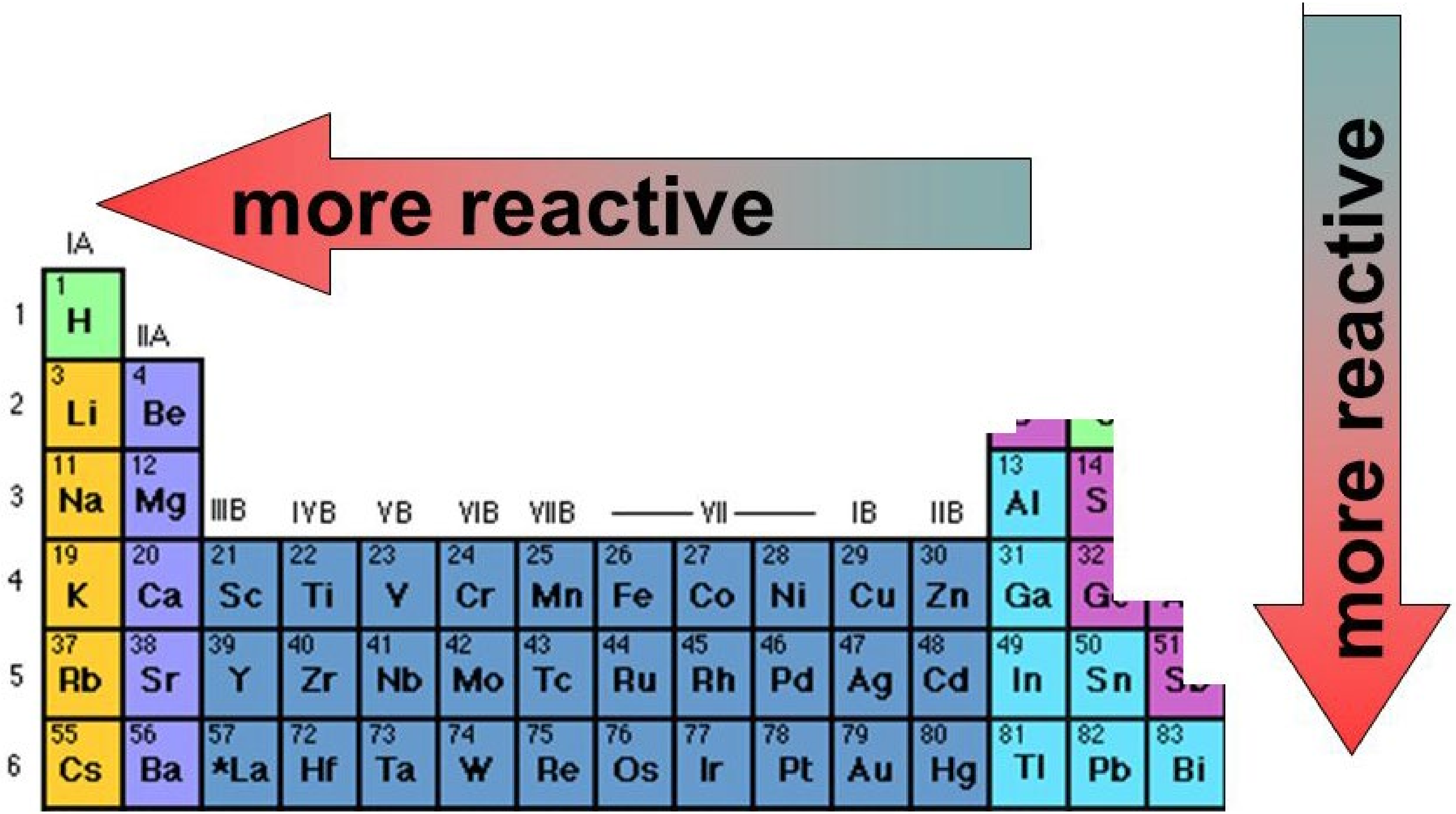
Trends in the PT

Property	Moving Down the Group	Moving Right across the Period
Atomic Radius	Increases	Decreases
Ionisation Energy	Decreases	Increases
Reactivity with Water	Increases	Decreases
Electronegativity	Decreases	Increases
Metallic Character	Increases	Decreases

Atomic Radius



Reactivity



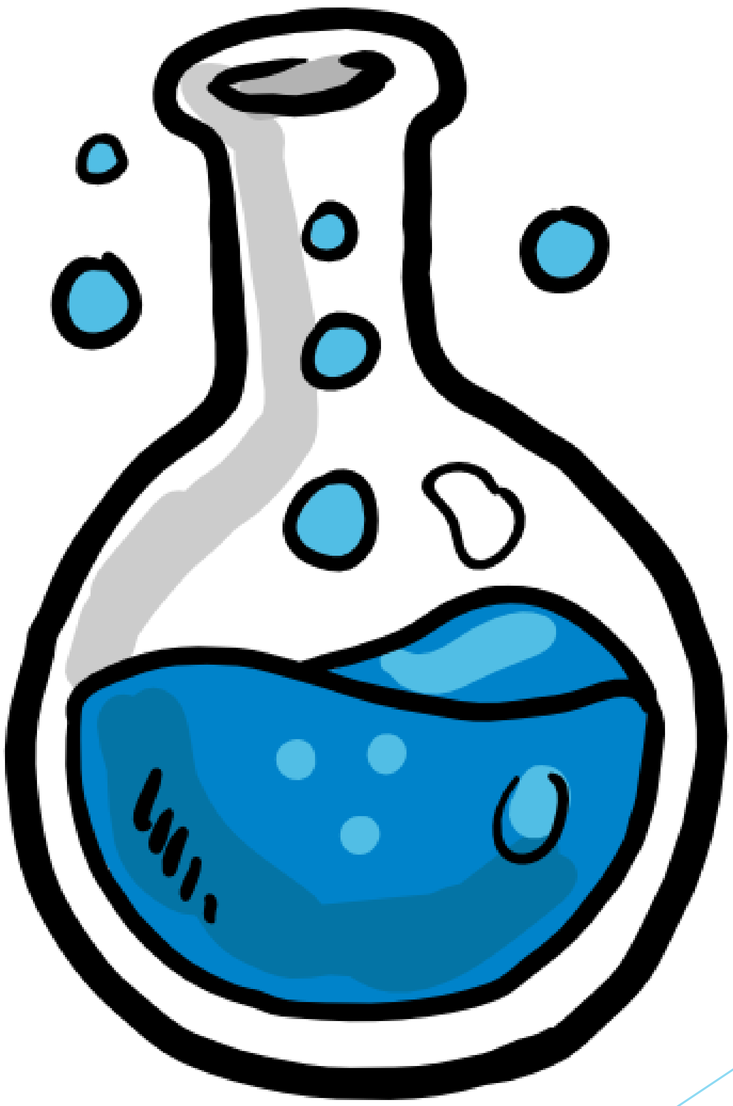
more reactive
more reactive
IA
IIA
IIIB
IVB
VB
VIB
VIIB
VII
IB
IIB
1 H
3 Li
4 Be
11 Na
12 Mg
13 Al
14 S
19 K
20 Ca
21 Sc
22 Ti
23 V
24 Cr
25 Mn
26 Fe
27 Co
28 Ni
29 Cu
30 Zn
31 Ga
32
37 Rb
38 Sr
39 Y
40 Zr
41 Nb
42 Mo
43 Tc
44 Ru
45 Rh
46 Pd
47 Ag
48 Cd
49 In
50 Sn
51
55 Cs
56 Ba
57 *La
72 Hf
73 Ta
74 W
75 Re
76 Os
77 Ir
78 Pt
79 Au
80 Hg
81 Tl
82 Pb
83 Bi
1
2
3
4
5
6

Electronegativity

0.5-1.0	1.0-1.5	1.5-2.0	2.0-2.5	2.5-3.0	> 3.0

1A	2A	3B	4B	5B	6B	7B		8B		1B	2B	3A	4A	5A	6A	7A	0
1 H 2.20																	2 He
3 Li 0.98	4 Be 1.57											5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98	10 Ne
11 Na 0.93	12 Mg 1.31											13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16	18 Ar
19 K 0.82	20 Ca 1.00	21 Sc 1.36	22 Ti 1.54	23 V 1.63	24 Cr 1.66	25 Mn 1.55	26 Fe 1.83	27 Co 1.88	28 Ni 1.91	29 Cu 1.90	30 Zn 1.65	31 Ga 1.81	32 Ge 2.01	33 As 2.18	34 Se 2.55	35 Br 2.96	36 Kr 3.00
37 Rb 0.82	38 Sr 0.95	39 Y 1.22	40 Zr 1.33	41 Nb 1.6	42 Mo 2.16	43 Tc 1.9	44 Ru 2.2	45 Rh 2.28	46 Pd 2.20	47 Ag 1.93	48 Cd 1.69	49 In 1.78	50 Sn 1.96	51 Sb 2.05	52 Te 2.1	53 I 2.66	54 Xe 2.60
55 Cs 0.79	56 Ba 0.89	71 Lu	72 Hf 1.3	73 Ta 1.5	74 W 2.36	75 Re 1.9	76 Os 2.2	77 Ir 2.2	78 Pt 2.28	79 Au 2.54	80 Hg 2.00	81 Tl 1.62	82 Pb 1.87	83 Bi 2.02	84 Po 2.0	85 At 2.2	86 Rn 2.2
87 Fr 0.7	88 Ra 0.9	103 Lr	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

Chemical Bonding

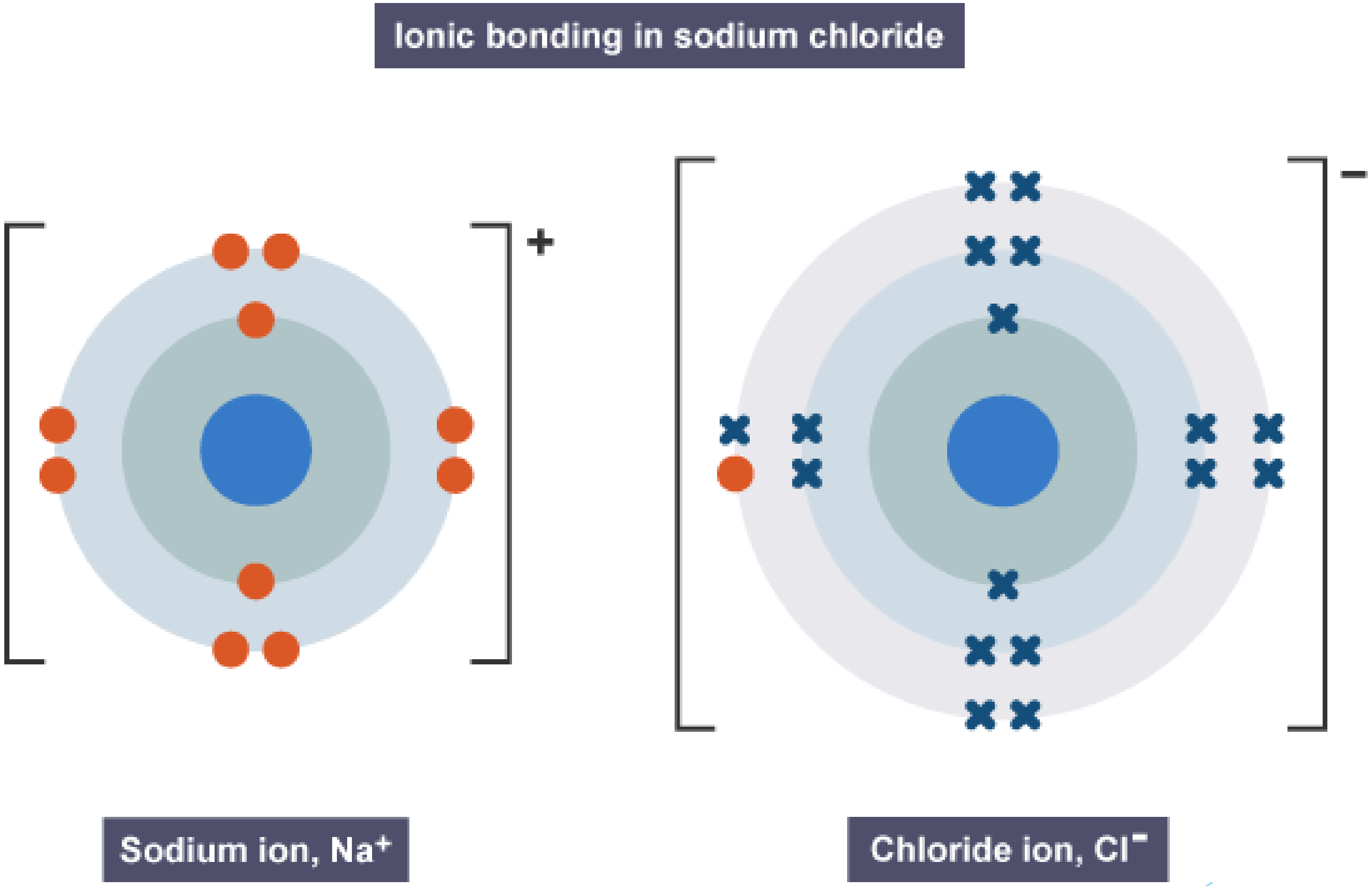


Ionic Bonding

- Atoms form **ions** – charged atoms formed by gaining or losing electrons
- Positive **Cations**, Negative **Anions**
- Ions have a **valency** – a charge calculated by the difference in protons and neutrons
- Opposite charges attract, sticking the atoms together in a lattice structure
- Ionic compounds are usually a metal and a non-metal, and are written as empirical formulas (indicating ratio)
- Naming ionic compounds follows the rule:

Metal Non-Metal-ide

Ionic Bonding



Polyatomic Ions

- The individual ion is made up of two or more atoms joined together
- Although made up of two or more atoms, each of these ions acts as a unified entity when forming compounds.

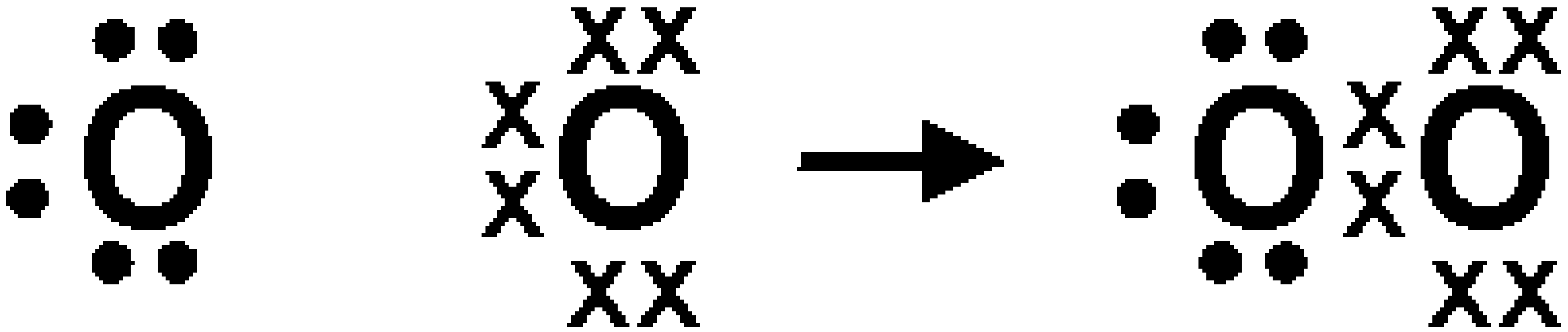
Polyatomic Ion	Formula	Ionic Formula	Charge
Ammonium	NH_4	$[NH_4]^+$	1+
Hydroxide	OH	$[OH]^-$	1-
Nitrate	NO_3	$[NO_3]^-$	1-
Sulfate	SO_4	$[SO_4]^{2-}$	2-
Carbonate	CO_3	$[CO_3]^{2-}$	2-
Phosphate	PO_4	$[PO_4]^{3-}$	3-

Covalent Bonding

- Covalent bonds are formed when electrons are shared to fill the outer shell of the atoms involved
- Covalent compounds have a molecular formula, specifying how many atoms of each substance within a molecule
- The valency of an element defines how many bonds it will form - indicated by the group number on the PT
- Usually formed between two non-metals, as there is not enough difference in electronegativity to force the movement of electrons from one atom to another

Example: Oxygen (O_2) molecule

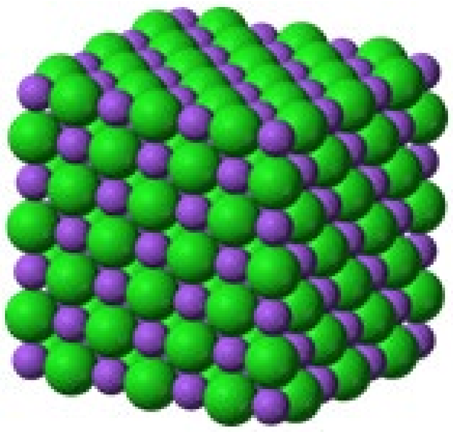
- Oxygen's ion is O^{2-}, therefore it will form two bonds when bonding covalently to fill its outer shell

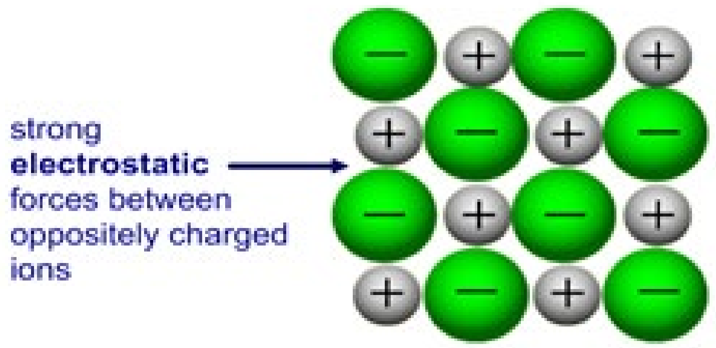


Properties of Ionic Substances

Electrostatic attraction between ions makes:

- Solids at room temperature
- High melting points
- Hard – lattice structure gives stability to each ion within it
- Do not conduct electricity as solids
- Brittle – distorting the crystal brings like chares together and the repel one another, causing the crystal to break



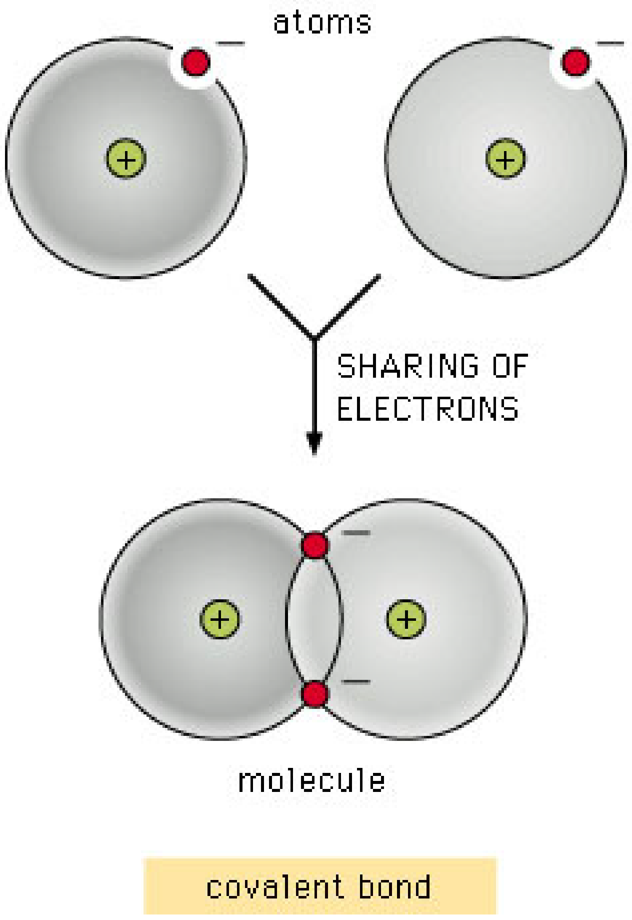


Properties of Covalent Substances

While bonding forces holding atoms together within covalent molecules are very strong, intermolecular forces (the forces between each molecule to another molecule) are much weaker.

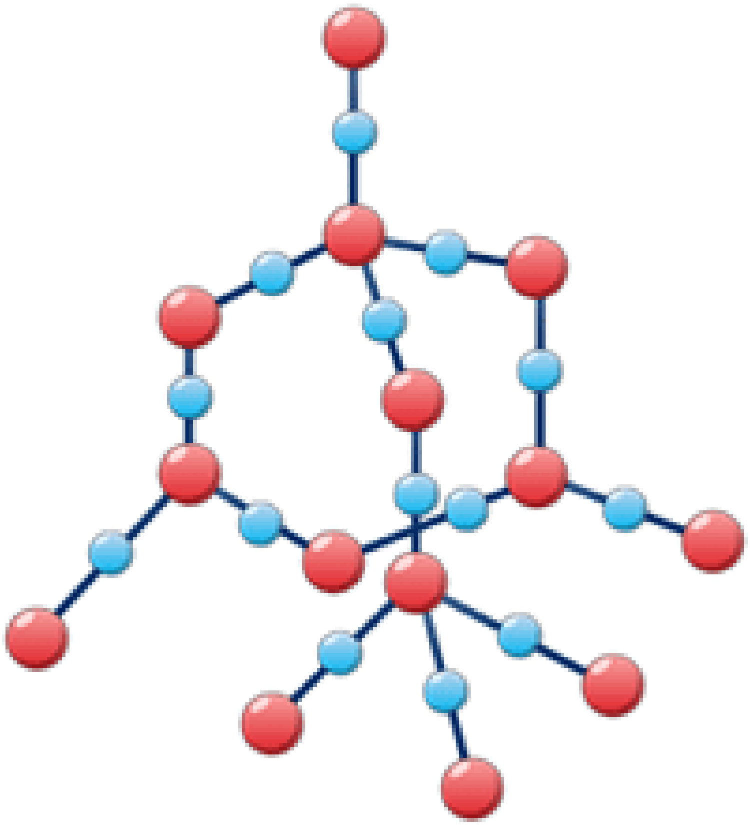
As a result, covalent substances are:

- Generally gases at room temperature
- Low melting points – less than 200°C
- Low boiling points – less than 400°C
- Soft solid forms
- Do not conduct electricity as solids or liquids



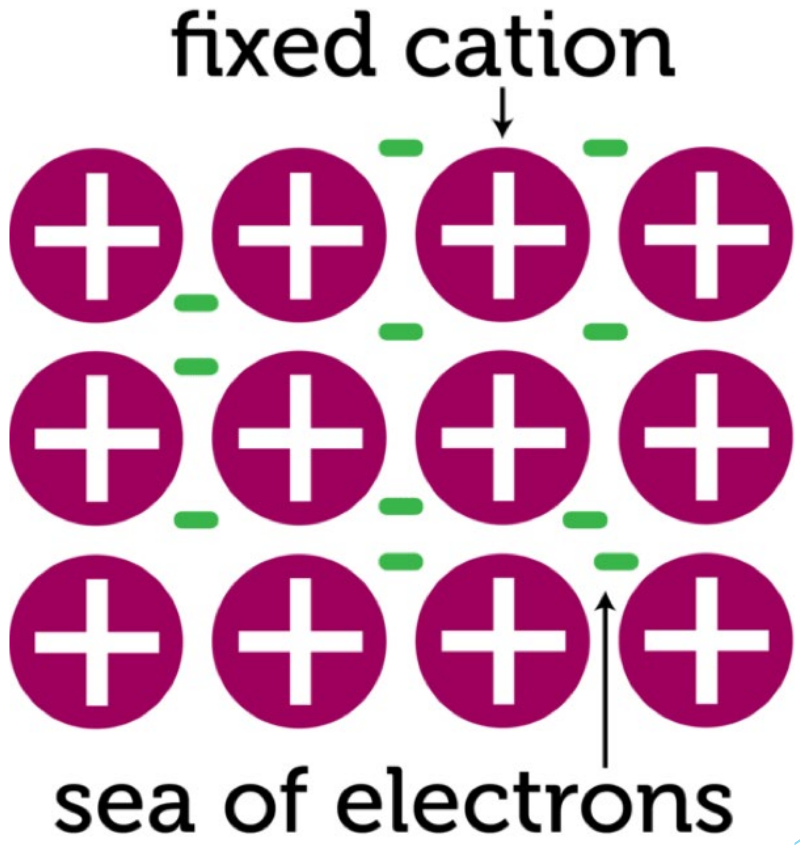
Covalent Network Solids

- Are solids where covalent bonding occurs indefinitely, hence have empirical formulae not molecular.
- Examples:
  - Silica (SiO_2) aka Quartz
  - Diamond (carbon)
- When covalent lattices melt, it involves breaking many covalent bonds that are very strong.
- This process requires a lot of energy, only occurring at very high temperatures (>1000°C)

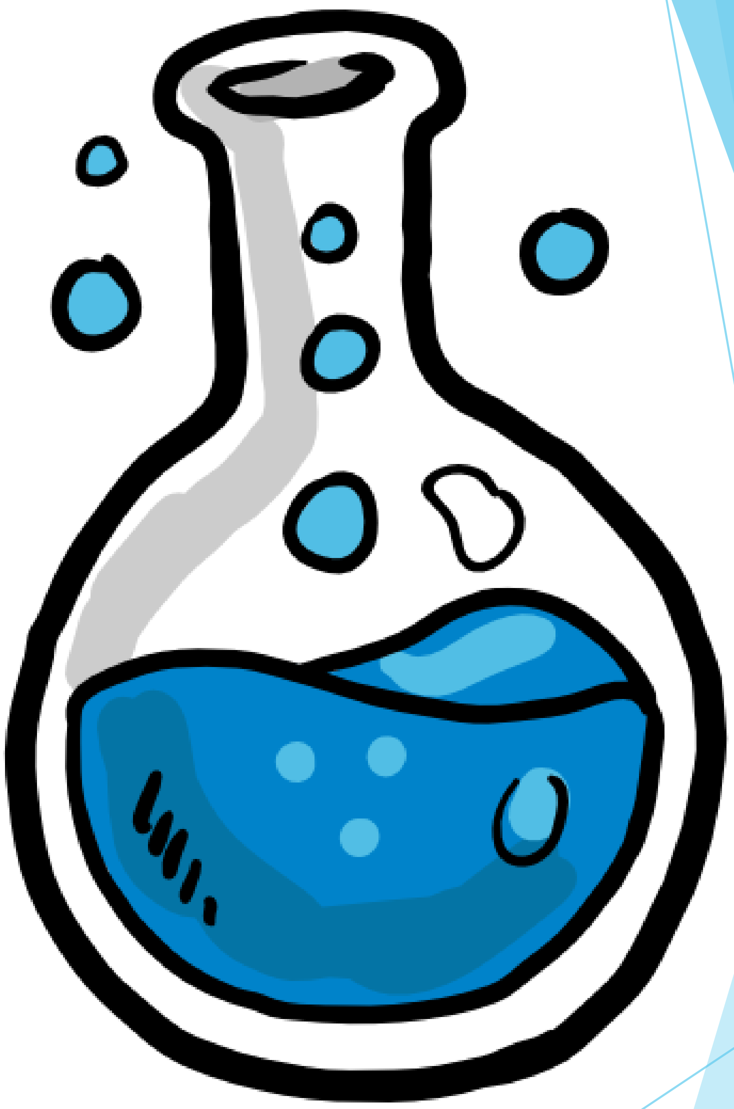


Metallic Bonding

- Positive ions immersed in a sea of electrons
- Gives high conductivity as electrons can flow through the structure
- Mobile electrons are able to hold the positive ions of metals together even when the metal is distorted giving malleability ad ductility of metals



Intermolecular Forces & Allotropy

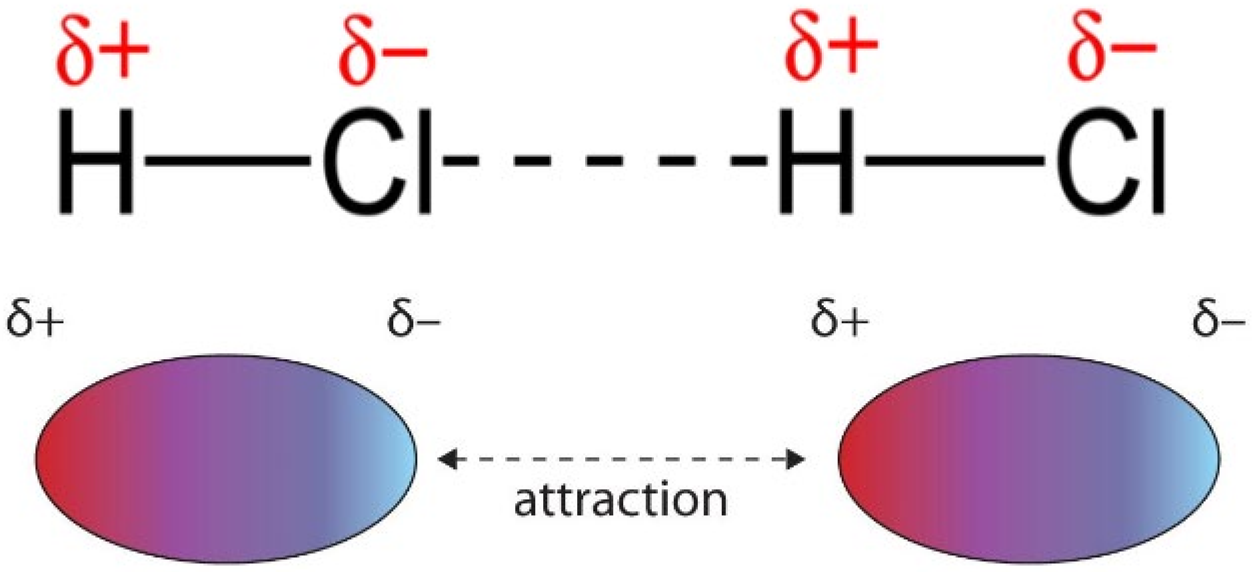


Electronegativity and Polarity

- If the two elements forming a covalent bond have different electronegativity's, the electrons being shared will be dispersed unevenly
- The electrons are attracted to the more electronegative atom, making the space around that atom slightly negative (electrons have a negative charge)
- As a result, the other atom in the bond has more space around it without electrons present, and as a result become slightly positive

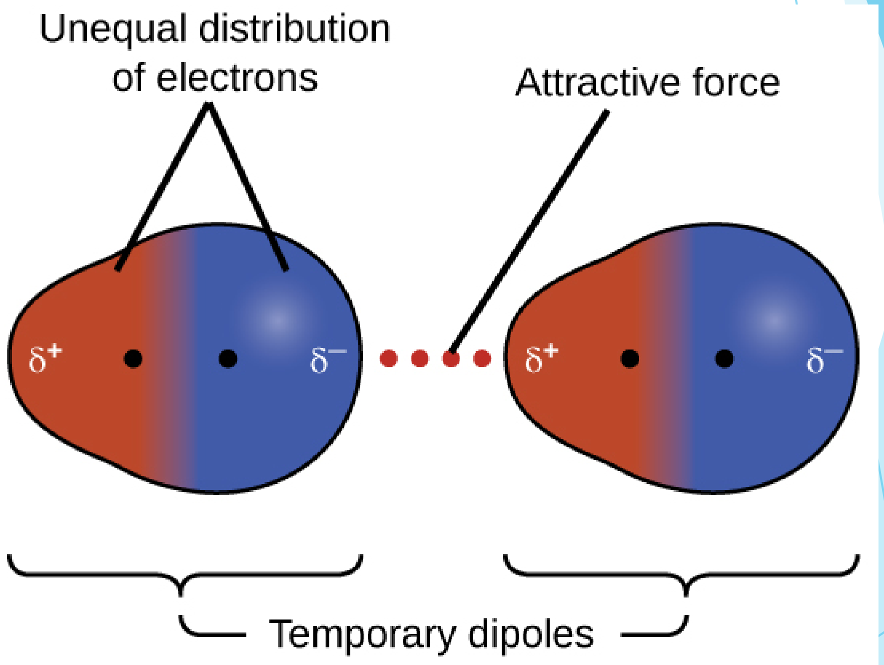
Dipole-Dipole

- Attractive forces that occur between polar molecules
- Opposite poles attract



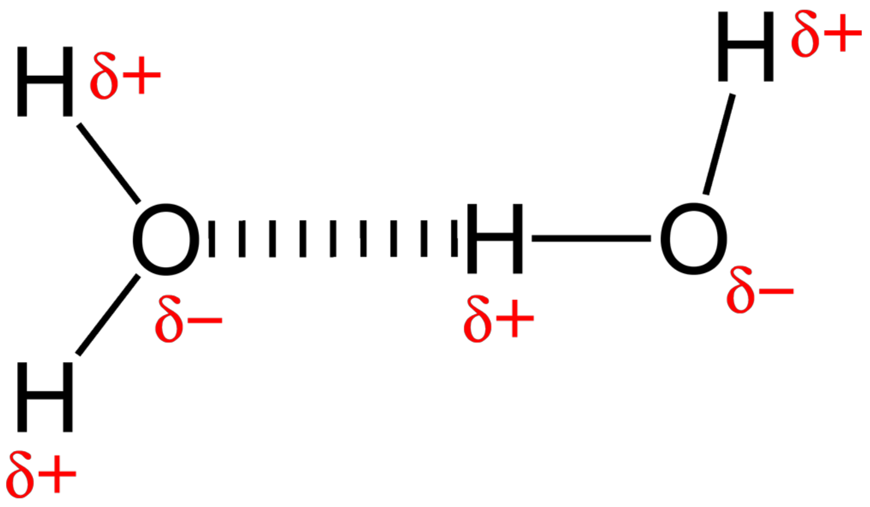
Dispersion Forces

- Dispersion forces are weak attraction forces that are electrostatic attractions between instantaneous dipoles in neighbouring molecules
- Because electrons are always in motion, at any particular time the electron cloud may be unevenly distributed, creating a tiny pole, attracting one molecule to the next



Hydrogen Bonding

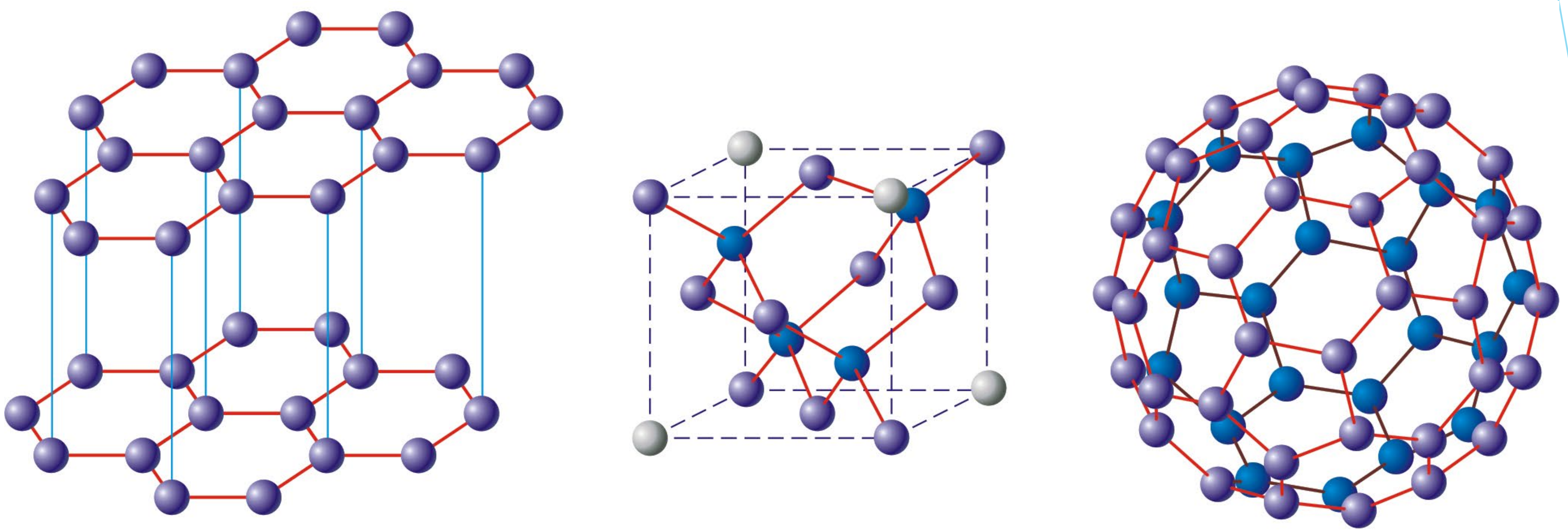
- Hydrogen is attracted to Oxygen, Nitrogen or Fluorine in another molecule, due to their high electronegativity.
- Within a molecule, the O, N of F holds the electrons tightly, becoming negatively charged and the hydrogen becomes positively charged
- This means that when two adjacent molecules come together, they have a strong electrostatic attraction, forming a strong intermolecular force



Allotropy

- Existence of two or more physically distinct forms of one element in the same physical state
- Allotropes show distinct differences in:
  - Colour
  - Density
  - Hardness
  - Electrical conductivity
- Differences in properties is resultant from the different structure or packing of atoms
- Typical allotropes include:
  - Phosphorous
  - Sulfur
  - Oxygen

Allotropes of Carbon



graphite diamond fullerene

MODULE 1 COMPLETE!