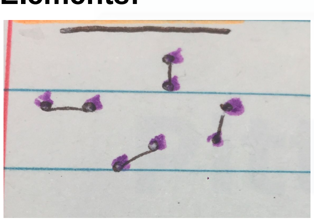
<u>Notes from Chem teacher PD</u>

- The syllabus has changed but the people who mark it are the same. And thus they will most likely mark it in a very similar way to previous exams
- Providing specific detail that showcases chemical understanding is critical to top performance band responses
- Spelling of chemicals is important. It has the spelling on the periodic table. If the value is in the data sheet, use it! NOT SOME OTHER SIMILAR VALUE
- There will be a clarification on organic naming that will be released soon by NESA rule of thumb for now though, is to stick to simple naming conventions. Most likely that if the name is ambiguous, it will not be asked
- If you are given a stimulus, use it in your answer
- Inquiry questions can be asked as exam questions. (expect one in the hsc as the long response (8 mark q)
- “Aboriginal question will be most likely be in the HSC” - my guess… you can bank on one of the questions being in the HSC. There are 2 in yr 12 course. (use of acids and bases in cycad nuts)
- Expect left field questions - use your chemical understanding and science skills (logical thinking, prac skills, data skills) to answer there questions “ how to I bring in my chemical knowledge”. Markers will expect you to write about something you have learnt
- Cross module question - 8 marker will probably ask you to bring in knowledge from many modules. Probably module 8 and other modules together. (analysis techniques of chemical compounds). E.g. environmental. E.g. easy exam questions to set. Examine the environmental, economic and sociocultural implications o obtaining and using hydrocarbons from the earth. (Env. pollution and land clearing, mining (what techniques could be used to monitor env impact Economic → good short term, long term need to transition energy and economy away from it. Sociocultural: what products can be produced that are beneficial / negative?

Module 1:

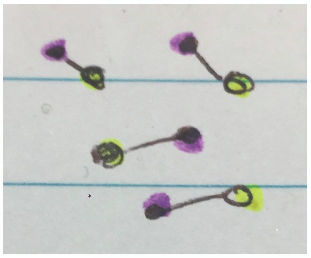
Elements:



-purest substances

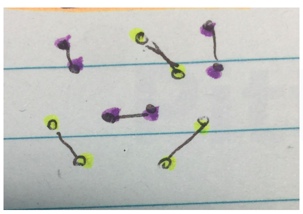
-made of only one type of atom

Compounds:



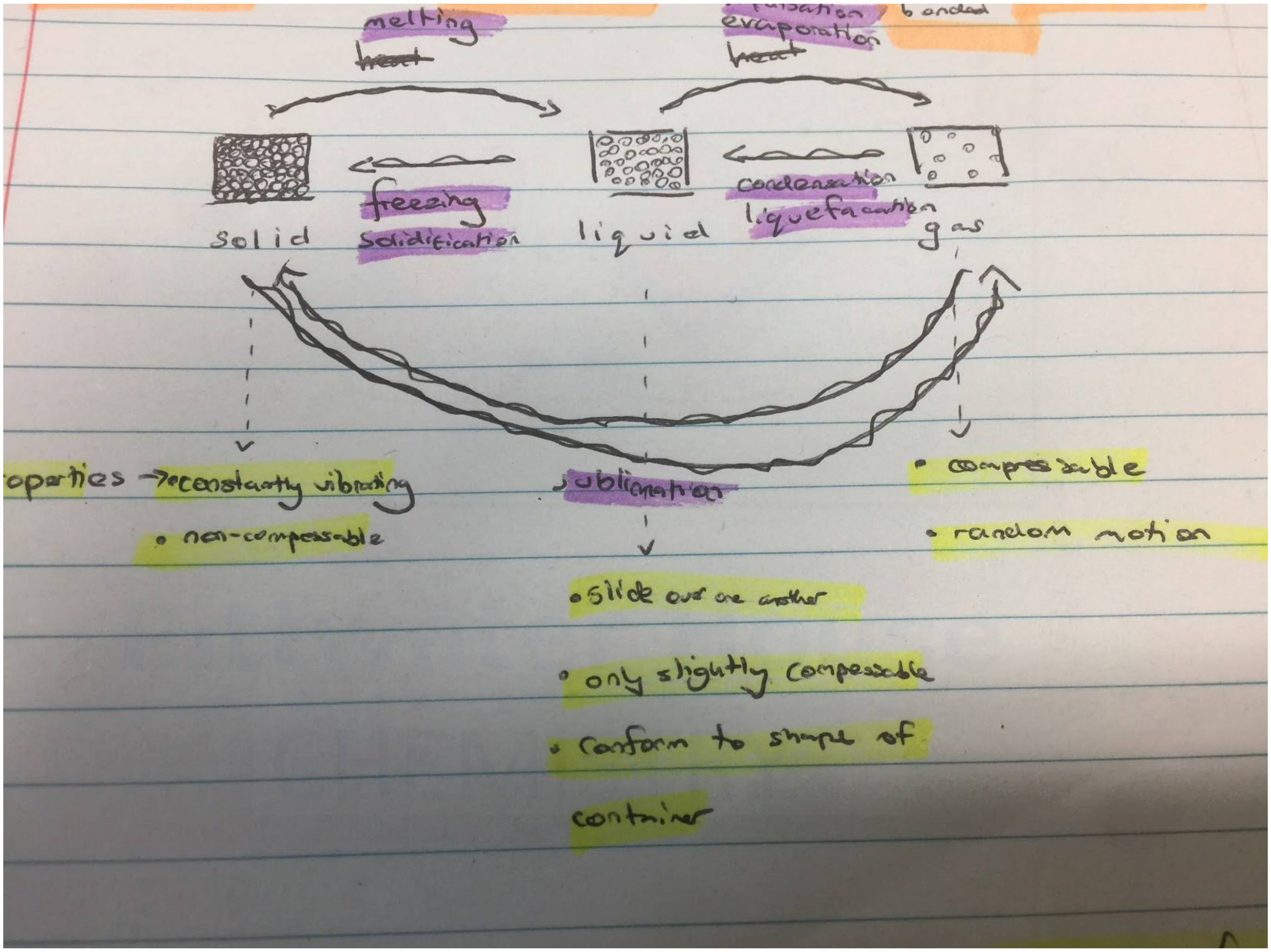
-more than one element chemically bonded together

Mixture:



-More than one element or compound together but not chemically bonded

Diagram of change



Synthesis -> creating a bigger compound from 2 or more smaller ones
Decomposition-> breaking apart a bigger compound into 2 or more smaller smaller ones

Homogeneous mixture -> same / uniform compositions -> eg. salt in water
Heterogeneous mixture -> different / separate (non uniform comp)-> eg. muddy water
Heterogeneous mixtures are generally opaque

Experiment	Dissolving sugar	Burning Magnesium	Dry ice disappears	Crush calcium carbonate	Dissolving ethanol	Hot copper carbonate

Observation	-Went clear in the water	- Burned very brightly -Changed colour	-frost forms because of water in the form of gas in the air	-Smaller particles	-Liquid dissolved in another liquid	-Change in colour -Letting off coloured gas
Chemical or physical change	Physical	Chemical	Physical	Physical	Physical	Chemical

Anion: negatively charged
Cathion: positively charged

Group	Metals	Semi-metals	Non-metals
Appearance	lustrous	Low sheen	dull
Electrical conductivity	high	Low (semiconductors)	None (insulators)
Thermal Conductivity	high	high	Low (insulators)
Malleability and ductility	high	moderate	None (brittle)
Density	Generally high	intermediate	low
Boiling point	Generally high	Very high	low
Strength	high	variable	low
Examples	Sodium, magnesium, iron, chromium, zinc, platinum, gold, mercury, lutetium	Boron, silicon, germanium, arsenic, antimony, tellurium, astating	Hydrogen, helium, carbon, nitrogen, oxygen, flouring, neon, phosphorus

Strontium - Metal - Firework Colouring
Barium - Metal - Firework Colouring
Lithium - Metal - Firework Colouring
Sodium - Metal - Firework Colouring
Silicon - Semi - Computer Chips
Arsenic - Semi - Paint/Wallpaper/A lot of things
Chlorine - Non - Bacteria killer for Pool
Sulfur - Non -

Solids of different size	Sieve
Solids in liquids	Filtration
Dissolved solids in liquids, dissolved liquids	Distillation / Evaporation / Decanting
Liquids	Fractional Distillation
Gases	Fractional Condensation (chromatography)
Different density	
Homogeneous	Mainly distillation
Heterogeneous	Filtration, Chromatography, Sieving, Centrifuge
Gas in liquid	

Property	Protons	Neutrons	Electrons
Charge	Positive	Neutral	Negative
Mass (Atomic Mass Units)	1	1	1/2000
Location	Nucleus	Nucleus	Electron Cloud (shell)
Symbol	**P** or **+**	**N**	**E** or **-**

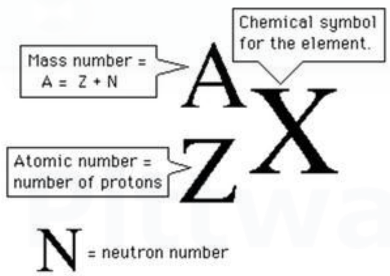
The proton determines what the element is

Atomic mass = the average mass of the naturally occurring isotopes

Nucleons (protons + neutrons) … aka. The number of particles in the nucleus

Isotope = same, number of protons, different neutrons

Eg. beryllium has 4 protons and 4 electrons so 5 neutrons



Name	Isotopic Notation	Number of Protons	Number of Neutrons
Potassium - 41	$_{19}^{41}K$	19	22
Magnesium - 26	$_{12}^{26}Mg$	12	14
Hafnium - 180	$_{72}^{180}Hf$	72	108
Uranium - 235	$_{92}^{235}U$	92	143

There are about 400 stable isotopes

Atoms of an element with the same atomic number but different mass number. (same protons but different neutrons
Outside the zone of stability atomic number $Z \leq 20$, proton:neutron ratio ≠ 1:1 (1:1 is stable)
Outside the zone of stability atomic number $21 < Z \leq 92$, proton:neutron ratio ≠ 1:1.5 (1:1.5 is stable)
Atomic number $Z > 92$ (transuranic) all are unstable

there are 3 types of decay

Radiation	Description	Charge	Penetrating Power

$_2^4a$ (alpha) $_2^4He$ He^{2+}	Helium nucleus (2 protons and 2 neutrons)	2+ (because no electrons)	Low
$_{-1}^0e$ (beta) (or $_1^0e$)	Electron (or positron, positive electron	1- (1+)	Medium
$_0^0\gamma$ (Gamma)	High energy electromagnetic wave in the X-ray region.	0 (neutral)	High

All are detectable with a Geiger counter/meter

Alpha decay of Uranium 238 - $_{92}^{238}U \rightarrow {}_2^4He + {}_{90}^{234}Th$

Alpha decay of Radon 226 - $_{88}^{226}Ra \rightarrow {}_2^4He + {}_{86}^{222}Rn$

Beta decay for:

$_{27}^{60}Co \rightarrow {}_{-1}^0e + {}_{28}^{60}Ni$
$_6^{14}C \rightarrow {}_{-1}^0e + {}_7^{14}N$
$_1^3H \rightarrow {}_{-1}^0e + {}_2^3He$

Some isotopes can be made by humans, others are naturally occurring.

How are they made?
Ion Accelerators - Two different lighter nuclei are accelerated towards each other at near speed of light speeds and crash together to form a new heavy nuclei

Neutron bombardment - Neutrons are fired into the nucleus of an atom and form new isotopes. This can push them outside the zone of stability and so they become radioactive.

Alpha bombardment - These are easier to speed up with electric fields as they are charged, and then can be crashed into a target nuclei, again, this can push the atom outside the zone of stability and cause it to become radioactive.

Process	Equation
Neutron Bombardment	
Alpha Bombardment	
Ion Acceleration	

Element - # ← Number (#) = nucleon number = proton + neutrons

(protons always stay the same)

Chlorine - 35 ← Number (35) = nucleon number = 17 + 18
(protons always stay the same)

Chlorine - 37 ← Number (37) = nucleon number = 17 + 20
(protons always stay the same)

Gamma radiation is an electromagnetic ray with very high energy
This radiation does not affect the atomic number of the atomic mass
Isotopes that decay via gamma (γ) radiation may have ***** or **M** to show it is radioactive (eg.: ${}_{27}^{60}Co^* \rightarrow \gamma + {}_{27}^{60}Co$

Radiation is able to penetrate different depending on its type: It is easiest to describe what each type of radiation can be stopped by.

Radiation type	Stopped by
Alpha	Paper
Beta	Thick aluminium, 0.5mm of lead
Gamma	5 cm of lead of thick concrete (15cm ish)

Half life:
- Unstable isotopes will radioactively decay into more stable isotopes.
- The rate of decay is determined by an isotopes half life. This is the amount of time it takes for half of the material to radioactively decay.
- Eg. if there was 50g of ^{42}K (${}_{19}^{42}K \rightarrow {}_{-1}^{0}e + {}_{20}^{42}Ca$) (beta emitter), and its half life is 2.6 days. How much radioactive ^{42}K is left after 5.2 days. 5.2 = 2 * 2.6 (2 half lifes), so after one half life there is 25g, after the second there is 12.5g.
- After each you half it again.

$Y = Ae^{-t\lambda}$

Y = amount of radioactive material after time
A = initial amount

T = time
λ = decay constant

Sodium - 2, 8, 1

Helium - 2

Nitrogen - 2, 5

Phosphorus - 2, 8, 5

Calcium - 2, 8, 8, 2

Bohr’s Electron Model
The model is good, it predicts valency and ionic charges of atoms. But the problem is that the model is only 2D, and in real life they are 3D.

Things we need to consider:
- Atoms are 3D not 2D, an atom’s electron cloud is 3D therefore
- What is the 3D shape exactly?
- Why do electrons seem to fit into these nice patterns?
- Quantum mechanics means that when tiny particles (electrons) are around a nucleus they take on characteristics of waves

Electrons around a nucleus are not particles but are “standing waves”.
Must be a whole number of waves, otherwise the waves break down. (it has to be 1 or 2 or 3 wavelengths, not 1.5) (deconstructive interference).
This is called the wave particle duality of matter (particles can be described as waves and waves can be described by matter).
Electrons around a nucleus are **not orbiting** like planets, but **are oscillating** in their standing waves around the nucleus.
Lower frequency waves have lower energy → high wavelength = low energy
De Broglie waves - each time you go up an energy you need to go up a whole number value of wavelengths - this corresponds to a specific amount of energy

N = 1 (1 wave)

N = 2 (2 waves)

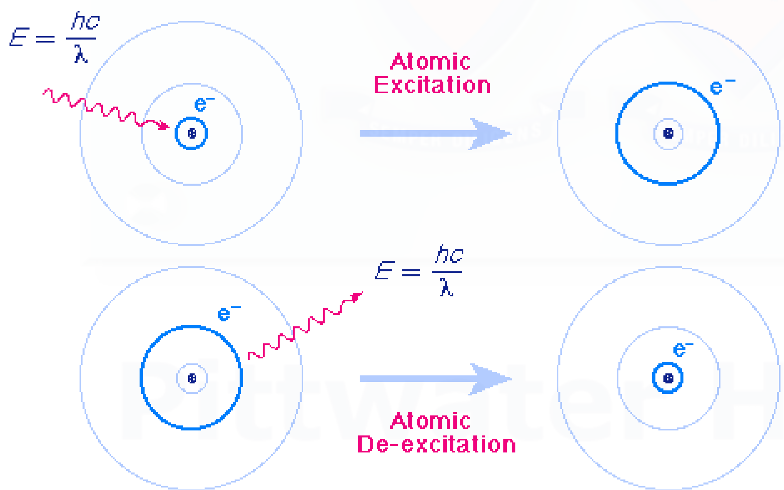
N = 3 (4 waves)

N = 4 (8 waves)

Some quantum basics - spectral lines
Quanta = small or discrete amount

Light can be absorbed or emitted by **electrons** in atoms at one set wavelengths that match the gap in energy that exists in the atom
This light emission can be observed as an emission spectrum. Before quantum mechanics this simple observation was not able to be explained.

Photon absorption



Photon emission

Notes on quantum mechanics

E = hf

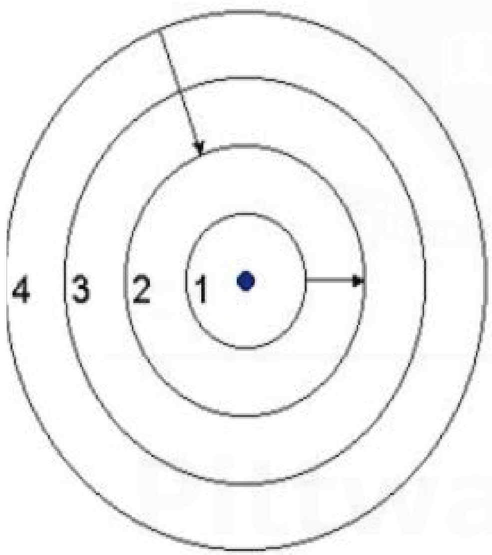
E = energy
H = planck's constant ($6.626*10^{-34}$ m^2 . kg . s^{-1})
F = frequency of light wave

Wave equation
$\lambda * f = v$
$E = hc/\lambda$

λ = wave length
F = frequency
V = velocity of wave (aka. Speed of light)

If λ is high, then E (energy) is lower. Eg. red light ~ 650 nm has a lower energy than violet light ~ 430 nm

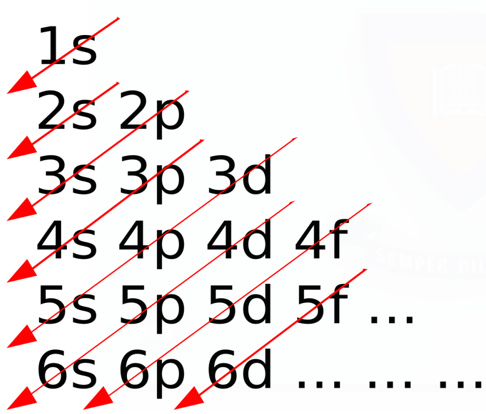
Different energy levels in atoms are seen here



A difference in energy level corresponds to a change in energy: ΔE

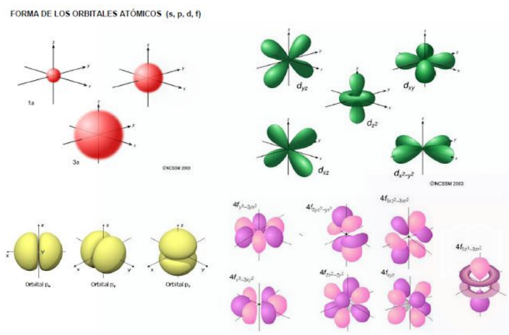
ΔE can be seen as light energy using a spectroscope. ΔE (change in energy of the atomic energy levels) is equal to the exact energy of the light emitted

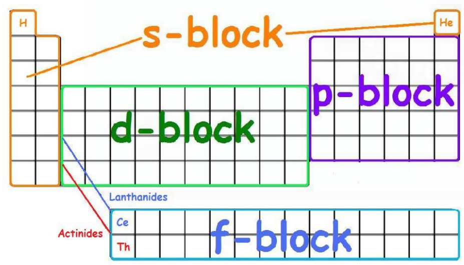
$\Delta E = hf$
$\Delta E = hc/\lambda$



- $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2$

In this diagram, the big number out the front of s, p, d or f is the principal quantum number
Each of these 1s or 3d hold 2 electrons
Each s orbital can hold 2 electrons
Each p orbital can hold 6 electrons
Each d orbital can hold 10 electrons
Each f orbital can hold 14 electrons





Eg.
C - 6e
$1s^2\ 2s^2\ 2p^2$
Mg - 12e
$1s^2\ 2s^2\ 2p^6\ 3s^2$

O - 8e
$1s^2\ 2s^2\ 2p^4$

Cl - 17e
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

Li - 3e
$1s^2\ 2s^1$

Fe - 26e
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$

Electron Configuration
- Notice patterns of electron configuration and how the periodic table is set out
- Elements in the same group will have the same valence configuration eg. Li ($1s^2\ 2s^1$) and Na ($1s^2\ 2s^2\ 2p^6\ 3s^1$) both have (s^1) on the end
- O ($1s^2\ 2s^2\ 2p^4$) and S ($1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$)

Atomic Radii

- Top to bottom - atomic radii gets bigger (greater number of energy levels)
- Left to right on one period - atomic radius gets smaller (kinda confusing perhaps?)
- Reason: as you move from the left to the right, the electrons fill in the same (or similar orbital) which will have a similar geometry. The nucleus gets more positively charged as you move to the right, therefore the electrostatic attraction increases pulling the spdf orbitals closer to the nucleus. Therefore the electrostatic attraction increases.

First Ionisation
- Amount of energy required to remove the valence electron
- $A \rightarrow A^{+} + e^{-}$
- Top to bottom - decrease in 1st ionisation
  + Reason: each period there is another shell of electron added so the outside electron must be further away
  + Outer electron is the one that is taken away from the atom (valence electron)
  + Electrostatic force between nucleus (+) and electron (-) decreases as distance increases
  + The further down the periodic table you go, the more thells there are in the atom and the further away from the nucleus the valence electron is. This means a weaker electrostatic force and so the amount of energy required to remove on electron decreases.
- Left to right - increase in energy
  + As you more along the period, the distance of the valence electrons from the nucleus decreases slightly and more importantly, the charge of the nucleus increases. This means the electrostatic force is relatively stronger as you more from left to right along the period. This makes the valence electrons harder to remove and you need more energy to take one electron away.

Atomic Number	Symbol	1st ionisation energy kJ/mol
1	H	1312
2	He	2372
3	Li	520
4	Be	900
5	B	801
6	C	1087
7	N	1402
8	O	1314
9	F	1681
10	Ne	2081
11	Na	496
12	Mg	738
13	Al	578
14	Si	787
15	P	1012
16	S	1000
17	Cl	1251
18	Ar	1521
19	K	419
20	Ca	590
21	Sc	633
22	Ti	659
23	V	651
24	Cr	653

25	Mn	717
26	Fe	763
27	Co	760
28	Ni	737
29	Ci	746
30	Zn	906
31	Ga	579
32	Ge	762
33	As	946
34	Se	941
35	Br	1140
36	Kr	1351

Electronegativity
- The ability for an atom to draw an electron to itself. Based on the electrostatic force of attraction of the (+) nucleus and the ability to create a full electron orbital.
- Left to right - increase
- Top to bottom - decrease
- Most electronegative element is FLUORINE
    - Some other electronegative elements: oxygen, chlorine, nitrogen

Group (vertical) / Period (horizontal)	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H 2.20																	He
2	Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne
3	Na 0.93	Mg 1.31											Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16	Ar
4	K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr 3.00
5	Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 1.9	Ru 2.2	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe 2.60
6	Cs 0.79	Ba 0.89	*	Hf 1.3	Ta 1.5	W 2.36	Re 1.9	Os 2.2	Ir 2.20	Pt 2.28	Au 2.54	Hg 2.00	Tl 1.62	Pb 2.33	Bi 2.02	Po 2.0	At 2.2	Rn 2.2
7	Fr 0.7	Ra 0.9	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Uub	Uut	Uuq	Uup	Uuh	Uus	Uuo
Lanthanides	*	La 1.1	Ce 1.12	Pr 1.13	Nd 1.14	Pm 1.13	Sm 1.17	Eu 1.2	Gd 1.2	Tb 1.1	Dy 1.22	Ho 1.23	Er 1.24	Tm 1.25	Yb 1.1	Lu 1.27		
Actinides	**	Ac 1.1	Th 1.3	Pa 1.5	U 1.38	Np 1.36	Pu 1.28	Am 1.13	Cm 1.28	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr 1.291		

Periodic table of electronegativity using the Pauling scale

Bonding (difference in electronegativity determines bond type)
- **Ionic** - Ionic electron has left the first atom and joined the second atom because its pull is so strong and the first's is so weak
    - (*3.3 - 2.1 = ionic*) e.g. NaCl 3.16 - 0.93 = 2.23
    - Ionic bonds occur between oppositely charged ions (often metal and non-metal but not always)
    - Strong electrostatic force holding them together (+) - (-)
    - Form regular crystal lattice structures
    - Strong bonds
- **Polar covalent** - when there is a moderate difference in electronegativity but the weaker atom does not completely let go of the electron
    - (*2.1-0.5 = polar covalent*) HCL 3.16 - 2.2 = 0.96
- **Nonpolar covalent** - both atoms have a fairly even pull on the electron, so it sits evenly in the middle of the nuclei
    - (*0.5 - 0 = nonpolar covalent*) C-H (carbon hydrogen bond) 2.55 - 2.2 = 0.35
- All diatomic gases have non-polar bond e.g. H-H, 0=0

Na-Br	-	Ionic	-	| 0.93 - 2.96 |	= 2.03
Mg-Cl	-	Polar Covalent	-	| 1.31 - 3.16 |	= 1.85
O-C	-	Polar Covalent	-	| 3.44 - 2.55 |	= 0.89
Ca-O	-	Ionic	-	| 0.11 - 3.44 |	= 2.44
H-F	-	Polar Covalent	-		
O-N	-	Non polar covalent	-		
C-N	-	Non polar covalent	-		

Lewis dot diagram for ions - square brackets with no dots for +ve ions, 8 dots for -ve ions

- They only include the symbol and the valence or outer electrons
- Eg. magnesium ion (+)

- Eg. Iodine ion (-)

Name	Formula
Carbonate	CO_3^{2-}
Nitrate	NO_3^-
Sulfate	SO_4^{2-}
Phosphate	PO_4^{3-}
Hydroxide	OH^-
Ammonium	NH_4^+
Nitrite	NO_2^-
Sulfite	SO_3^{2-}
Acetate	CH_3COO^-

Name	Formula
Water	H_2O
Methane	CH_4
Carbon Dioxide	CO_2
Ammonia	NH_3
Hydrogen chloride (hydrochloric acid)	HCl

Valency - bonding power

- Valency of 1 corresponds to its ability to make one bond
- Valency of 2 = 2 bonds

Covalent bonds

- For example water has 2 hydrogen atoms and 1 hydrogen atom. They are able to share their outer electrons and get full valence shells.
- Hydrogen has 1 and needs one more (its outer shell fills up to 2)
- Oxygen has 6 and needs 2 more (its outer shell fills up to 8)

Naming covalent compounds

- General naming rules the first elements in the formula generally appears in the first part of the name. The second element, the second part.
- The exact name depends on how many atoms of each element there are, e.g. SO_2 - the small 2 after the O means there are 2 oxygen atoms. The compound will therefore be a **di**oxide. The same of SO_2 is therefore sulfur dioxide.
- Mon usually only used for oxygen, e.g. CO → carbon monoxide

Number of Atoms	(1)	2	3	4	5	6
Prefix	(mon)	di	tri	tetra	penta	hexa

Types of structures:

- Ionic lattice (NaCl)
  - Positive and negative electrons for electrostatic attraction (opposite charges attract)
  - Form a regular arrangement. NaCl is a good example of an ionic lattice
  - Components of the lattice are the ions Na^+ and Cl^-

- Very strong bonds
- If they dissolve, the solution becomes conductive
- To find out the formula, use the ‘cross method’ aka. You cross the charges to find out how many of each element you need, so Sodium Bromide, Na is (+) charged and Br is (-) charged, so you need 1 br and 1 na to make a neutral compound

- Covalent molecular lattice (water ice, sugar)
    - Atoms share electrons to form covalent bonds. Covalent compounds formed are discrete units e.g. water, H_2, NH_3, CH_4, etc. **these are molecules**
    - Each unit in the lattice is a molecule
    - Moderate strength bonds - moderate strength attraction between lattice units
    - **Intermolecular forces** keep molecules together
    - Can form a lattice in solid structure
    - Non-conductive - good insulators no free electrons or ions from one molecule to the next.
- Covalent network lattice (diamonds)
    - Very strong bonds
    - Good insulators (e) - no free electrons or ions
    - Good insulators (heat) - vibrations are not passed through the lattice easily due to rigid bonds
    - Hard and brittle
    - Diamonds, silicon dioxide (sand)
    - Graphite (special case)
- Metallic lattice (iron)
    - Stationary metal ions
    - Valence electrons exist in a “sea of delocalised electrons” - they aren't stuck to one atom they are shared amongst neighbouring atoms
    - Electrostatic attraction between metal ion and sea of electrons created the metallic bonds
    - Metal ions occupy the units of the lattice
    - 3D lattice with a sea of electrons around it (electrons are + and the “sea” is negative, so overall the charge is neutral. (this means that you get an electrostatic attraction)
    - Free electrons allow electrical conductivity
    - Metal ions are easy to vibrate and increase movement therefore easy to conduct heat
    - The metallic bond is strong in general so high boiling/melting point
    - When metals are put under stress they are **malleable** and **bend** rather than being brittle. This is because metallic bonds can easily move and create new ones. They don't break but just move.
- Lattice: regular arrangement of particles (i.e. **crystal** lattice)

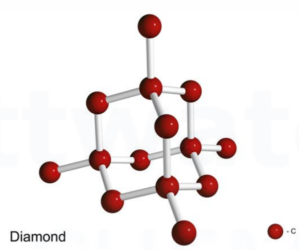
Use the diagrams to explain the differences between magnesium sulphide and magnesium in terms of

- malleability, and

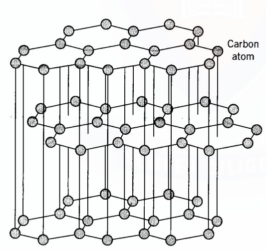
- Electrical conductivity

Allotropes

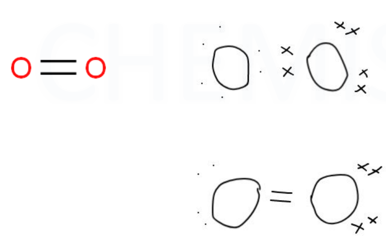
- Pure element
- 2 or more different structures
- Different structure → different properties
- **Carbon:**
    - Diamond
        - Covalent network
        - Tetrahedral covalent bonds going on “forever”
        - Very strong
        - Non-conductive (no free e)
        - Transparent
        - Used in drills and jewellery



    - Graphite
        - Covalent network
        - Flat sheets of carbon bonded in hexagons. Sheets can slide easily
        - E in-between layers
        - Electrical conductor
        - Brittle
        - Used as lubricant
        - Opaque



- **Oxygen:**
    - Oxygen Gas
        - Molecular structure
        - O_2
        - Non-polar, non-toxic gas, chemically reactive, fair oxidant

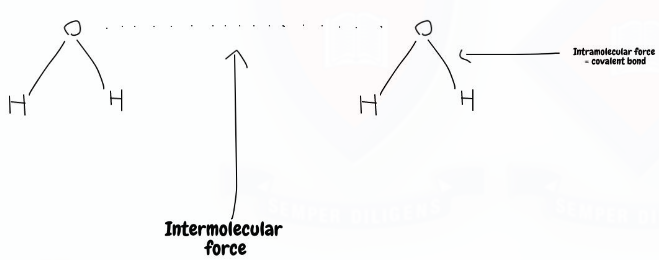


- Ozone Gas
    - Molecular structure
    - O_3
    - Polar, toxic gas, very chemically reactive, very good oxidant

With an ionic structure lattice between nonmetals and metals, the ionic lattice structure has alternating + and - ions in 3D that extend onto 'infinity'. These + and - ions are held together with an electrostatic force.
Non-conductive, except when molten or dissolved in water.
Rigid structure and brittle if struck, the ions move next to the - charge then repel and crack the crystal.

Inter/Intra molecular Forces:

- Intramolecular forces - the bonds inside a molecule. Either ionic or covalent
- Intermolecular forces - holds two of the same molecule together e.g. one water molecule attached to another water molecule



- Delta is used to show a little bit of charge (δ)
- Intermolecular forces affect the Boiling Point, Melting Point, Solubility, Viscosity, Surface Tension

Type of intermolecular force	Relative strength	Common?
Dispersion forces (van der waals forces)	Weak	All molecules - related to the size of the molecule
Dipole dipole	Moderate	Polar covalent molecules with overall dipole moment (the moment means the dipole has direction and is quantifiable with a vector)
Hydrogen bonds - they are not bonds	Strong	Polar covalent molecules with overall dipole moment when the bond is H-O, H-N or H-F They are a special type of dipole-dipole intermolecular force. Higher Melting point because hydrogen bonds are stronger than other types of intermolecular forces.

SUBSTANCE	BOILING POINT
H_2	-253
CH_4	-162
NH_3	-33
H_2O	100

Module 2:

Reactions

- Word Equation
- Chemical Formulas

- Balanced Chemical Equation
- Reactant → Product

Mole Equations

$n = m/MM$

n = chemical amount in moles
m = mass in grams
MM = molecular mass in g/mol (add all atomic masses)

$n * N_a = N$

n = moles
N_a = avogadro's number ($6.022 * 10^{23}$)
N = number of particles:
- Molecules
- Atoms
- Units of Ionic compounds

Example

	2Mg(s) +	O_2(g) →	2MgO(s)
Mole Ratio	2	1	2
Mole Mass	24.31	16*2 = 32	40.31 (24.31 + 16)
Mass	**5.52**	Moles*MM = 3.68	Moles*MM = 9.27
Moles	mass/MM = 0.23	mass/MM= 0.115	mass/MM = 0.23

%Component Composition = Component Mass / Molecule Mass

a) Find the percent mass composition of hydrogen and oxygen in water. H_2O.
%H = (M [H] = 2 x 1.008) / (M [H_2O] = 2 x 1.008 + 16) = 2.016 / 18.016 = 11%

%O = (M [O] = 16.00) / (M [H_2O] = 2 x 1.008 + 16) = 16.00 / 18.016

b) If you have 10g of water how much mass does the hydrigen contribure
H = 11% of mass when 11/100 * 10 = 1.1g

NaCl molecular formula would be $Na_{1000000000000000000}Cl_{10000000000000000000etc}$ but the simplest ratio of the ions is 1:1 therefore its formula is NaCl

The empirical formula of a substance is CH and its molecular mass is 78. Determine the molecular formula.
Take the empirical formula CH, we need to find C_b H_b so that the formula mass is = 78
(c) 12.01 x b + (h) 1.008 x b = 78, solve for b
OR
12.01 + 1.008 = 13.018
13.018 x ? = 78
? = 6 molecular formula = C_6H_6

$2Mg + O_2 \rightarrow 2MgO$

Materials

	Mass (g)	Moles
Crucible		
Magnesium		
Magnesium Oxide		
Oxygen		

Moles of Mg: moles of O

$Mg_{\square}O_{\square}$

Balanced chemical formula	$C_{12}H_{22}O_{11}$	12C	$11H_2O$
Mole ratio	1	12	11
Molar mass	342.296	12.01	18.016
Mass	11	4.64	6.36
Moles	0.03	0.386	0.353

Limiting Reagent

- What if there is a limit to the amount of reactants?
- $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$
- If there are 3 moles of HCl and5 moles of NaOH determine the moles of NaCl produced in the reaction?
- HCl gets used up first - we call this the limiting reagent
- The reactant that gets used up first is called the limiting reagent, the reactant that has a remainder we say is in excess
- HCl - limiting
- NaOH - excess

$HCl + NaOH \rightarrow NaCl + H_2O$

Balanced chemical formula	$HCl_{(aq)}$	$NaOH_{(aq)}$	$NaCl_{(aq)}$	$H_2O_{(l)}$
Mole ratio	1	1	1	1
Molar mass (from periodic table)	36.458	39.998	58.44	18.016
Mass (usually given)	5.00	5.00	7.305	2.252
Moles = mass / molar mass (use moles of limiting reagent to	0.137	0.125	0.125	0.125

determine moles of product)				

Balanced chemical formula	Mg	2HCl	H_2	$MgCl_2$
Mole ratio	1	2	1	1
Molar mass (from periodic table)	24.31	36.458	2.016	95.21
Mass (usually given)	5.3	9.5	0.263	12.42
Moles = mass / molar mass (use moles of limiting reagent to determine moles of product)	0.218 Only can use 0.1305 moles of Mg because 2HCl is 2 mole ratio and there is not enough HCl to have Mg x 1 and HCl x 2, we have to have HCl x 1 and Mg x 1/2	0.261	0.1305	0.1305

Concentration

Common units

- g/L

- mg/L
- g/g (w/w%)
- g/kg
- g/100g
- Ppm (dilute solutions, or gases)
- mol/L

- Defined as: relative amount of solute in a solution. E.g. how much salt in salt water (g/L), how much caffeine in coffee (mg/L), sugar in coke g/100mL.
- Molarity = moles of solute / volume of solution
- **C = n / V**
- C = concentration (molarity) mol/L $mol.L^{-1}$
- n = moles
- V = volume of solution in litres L

- Find the molarity of a solution containing 0.5 mol of sodium chloride in a 2L solution
- C = 0.5 / 2
- C = $0.25mol.L^{-1}$

When $2H_2S_{(g)}$ reacts with $3O_{2(g)}$ produces $2SO_{2(g)}$ and $2H_2O_{(g)}$
How many moles of SO_2 are formed and how many moles of reactant left?

a) 3 mol of H_2S and 4 mol of O_2

	$2H_2S_{(g)}$ +	$3O_{2(g)}$ →	$2SO_{2(g)}$ +	$2H_2O_{(g)}$
Mol Ratio	2	3	2	2
Given	3	4	(to get this we times by 4/3 as that is the 'ratio' of the limiting [from B5 and C5]) = 2.7	
Check H_2S limit	3 (from cell above)	3/2 * 3 (given / mol ratio of other reactant times mol ratio of this reactant) = 4.5 (we do not have 4.5 mol of O_2		
Check O_2 limit	4/3 * 2 (given / mol ratio of other reactant times	4 (From given)		

	mol ratio of this reactant) = 2.7			

Standard Solutions

- Standard solution: a solution of an accurately known concentration made from pure starting materials. Usually a solid salt and distilled water

{Good Substance}

- Properties necessary for salt used in standard solution:
  - Substance in highly pure solid form
  - Relatively high formula mass (minimise relative errors in weighing mass)
  - Soluble in water
  - Not react with air moisture
  - Good examples:
    - **a base**: anhydrous sodium carbonate (Na_2CO_3),
    - **an acid**: oxalic acid dihydrate ($C_2O_4H_2.2H_2O$)

{Bad Substance}

What makes a bad standard substance

- Absorbs moisture from the atmosphere
- Reacts with CO_2 in the air
- Low solid purity
- Is a gas at room temperature
- Is volatile and will evaporate
- **Bad bases - take moisture from the air, react with CO_2**
  - Potassium hydroxide (KOH)
  - Sodium hydroxide (NaOH)
- **Bad Acids - Volatile at room temp or gases**
  - HCl
  - H_2SO_4
  - HNO_3

How do you get pure, accurate solutions of NaOH and HCl common base and acid

- If you want to know sodium hydroxide accurately (NaOH) you can make a standard acid solution and do titration. What standard?

- If you want to know hydrochloric acid accurately (HCl) you can make a standard solution of base and titrate. What standard?
- Titration = quantitative technique used to accurately find the concentration of a solution
- (generally you will know roughly what the conc is, e.g. 0.1M HCl, but titration might reveal the conc to be 0.1011M. - This is much more accurate)

Dilutions

- Take a solution of an accurately known volume of an accurately known concentration solution and add a set amount of distilled water to ***decrease the concentration.***

  1. Create original solution. Collect and accurately known volume using a quantitative pipette
  2. Transfer this volume of solution to a volumetric flask
  3. Fill up the solution to the mark on the volumetric flask

{Diluting Problems($c_oV_o = c_dV_d$)}

- Use the equation: $c_oV_o = c_dV_d$
  - c_o - concentration of original solution
  - V_o - volume of original solution transferred
  - c_d - concentration of diluted solution
  - V_d - final volume of diluted solution

A 15mL of solution of 1.5M NaOH was diluted to make up a final volume of 250mL. Calculate the final concentration.

$1.5 \times 0.015 = c_d \times 0.250$, $c_d = (1.5 \times 0.015) / 0.250$, $c_d = 0.090\ mol.L^{-1}$

Thermodynamics + 4 main gas laws

Properties of Matter

Solids	Liquids	Gasses
Constantly vibrating Non-compressible Dense Fixed volume Fixed shape	Flow over each other Only slightly compressible Conform to shape of container Viscous	Compressible Random motion Pressure (kPa), (atm) Volume (L), (dm^3) Moles Temperature (°C), (K)

- P = pressure, V = volume, T = temperature, k = constant, n = moles of gas
- Gay-Lussac's Law - PV/T = k
- Boyle's Law - $P_1V_1 = P_2V_2$, OR PV = k, volume and pressure of a gas is changed. If one is changed, the other changes in proportion.
- Charles' Law - V = kT, volume of gas is related to its temperature
- Avogadro's Law, V = nk
    - At a set temperature and pressure, 1 mole of ANY gas takes up a set volume. In particular:
        - At 25°C and 100 kPa (1 atm) 1 mole of gas takes up 24.79L
        - V = nk, V = 24.79, n = 1, find k.
        - V/n = k, 24.79/1 = 24.79 (for 25°C, 100 kPa)
- Gay-Lussac's Law of combining volumes
    - "When measured at a constant temperature and pressure, the volume of gases taking part in chemical reactions show simple whole number ratios to one another"
    - E.g. 200 mL of hydrogen gas reacts with 100 mL of oxygen gas and created 200mL of water vapour (2:1:2)
    - (mole ratios are simple whole numbers) historically this was found out with gas reactions first because their volumes were easily recorded

Module 3:

<u>Synthesis</u>

- A chemical reaction where two or more chemicals join to make a larger one

<u>Decomposition</u>

- A large chemical breaks down into two or more chemicals. These reactions generally need some kind of catalyst of increase in heat to get them to break down

<u>Combustion</u>

- Involve the reaction between hydrocarbon and oxygen. These reactions are exothermic, they always let off energy.
- Hydrocarbon = molecule containing hydrogen and carbon

Precipitation Reactions

Precipitation Reaction:

- When 2 solutions (soluble salt + water) are mixed together and form an insoluble product (salt – ionic compound)

àPb(NO3)2(aq) + KI(aq) à PbI2(s) + KNO3()

àNaCl(aq) à Na+(aq) + Cl-(aq)

Complete Ionic equation: all dissolved species are separated
àSplit up all the (aq) ions
àThere shouldn't be any subscripts except for in **polyatomic ions**

Net ionic: only includes ions that change state or change chemically
(aq) + (aq) à (s)

- if it is identical on the left and right it can be cancelled

ions that are cancelled are called spectator ions
e.g. 2Cl à 2Cl (Cross them out)

Example:
2AgNO3(aq) + K2SO4(aq) à Ag2SO4(s) + 2KNO3(aq)

Full ionic equation: (all dissolved ions split up)
2Ag+(aq) + 2NO3_(aq) + 2K+(aq) + SO42-(aq) à Ag2SO4(s) + 2K+(aq) + 2NO3-(aq)

Net ionic equation: (only reactive species)
2Ag+ + SO42- à Ag2SO4

Complete formula equation

$Fe(NO_2)_{2(aq)} + 2NaOH_{(aq)} \rightarrow Fe(OH)_{2(s)} + 2NaNo_{3(aq)}$

- Full ionic
  - $Fe^{2+}_{(aq)} + 2NO_3^-{}_{(aq)} + 2Na^+_{(aq)} + 2OH_{(aq)} \rightarrow Fe(OH)_{2(s)} + 2Na^+_{(aq)} + 2NO_3^-{}_{(aq)}$
- Net Ionic (cancel out)
  - $Fe^{2+}_{(aq)} + 2OH_{(aq)} \rightarrow Fe(OH)_{2(s)}$

Acid Reactions

Acid + metal → salt + hydrogen
Acid + carbonate → salt + carbon dioxide + water
Acid + base → salt + water

Common acids: HCl, H_2SO_4, HNO_3, CH_3COOH (acetic acid - hard eg)
Common base: NaOH, KOH, NH_3 (ammonia hard eg, does not form water in acid base reactions)
All considered a type of neutralisation reaction - take away acidic properties of the acid

Acid Dissociation

Dissociation: Dissociation in chemistry is a process in which a molecule separates and splits into smaller parts such as ions
Acids dissociate in water to form hydrogen ions. What happens to each acid in water? Write a chemical equation.

$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$
$HNO_{3(aq)} \rightarrow H^+_{(aq)} + NO_3^-{}_{(aq)}$
$H_2SO_{4(aq)} \rightarrow 2H^+_{(aq)} + SO_4^{2-}{}_{(aq)}$

Hydrochloric acid + magnesium → Magnesium Chloride + Hydrogen
(hydrochloric acid makes chloride salts)
$2HCl_{(aq)}\ Mg_{(s)} \rightarrow MgC_2l_{(aq)} + H_{2(aq)}$

Sulfuric acid + sodium carbonate → Sodium sulfate + oxygen + water
(Sulfuric acid makes sulfate salts)
$H_2SO_{4(aq)} + Na_2CO_{3(s)} \rightarrow Na_2SO_{4(aq)} + CO_{2(g)} + H_2O_{(l)}$

Nitric acid + potassium hydroxide → Potassium nitrate + water
(nitric acid makes nitrate salts)
$HNO_{3(aq)} + KOH_{(aq)} \rightarrow KNO_{3(aq)} + H_2O_{(l)}$

Metal reactivity based on types on types of reactions

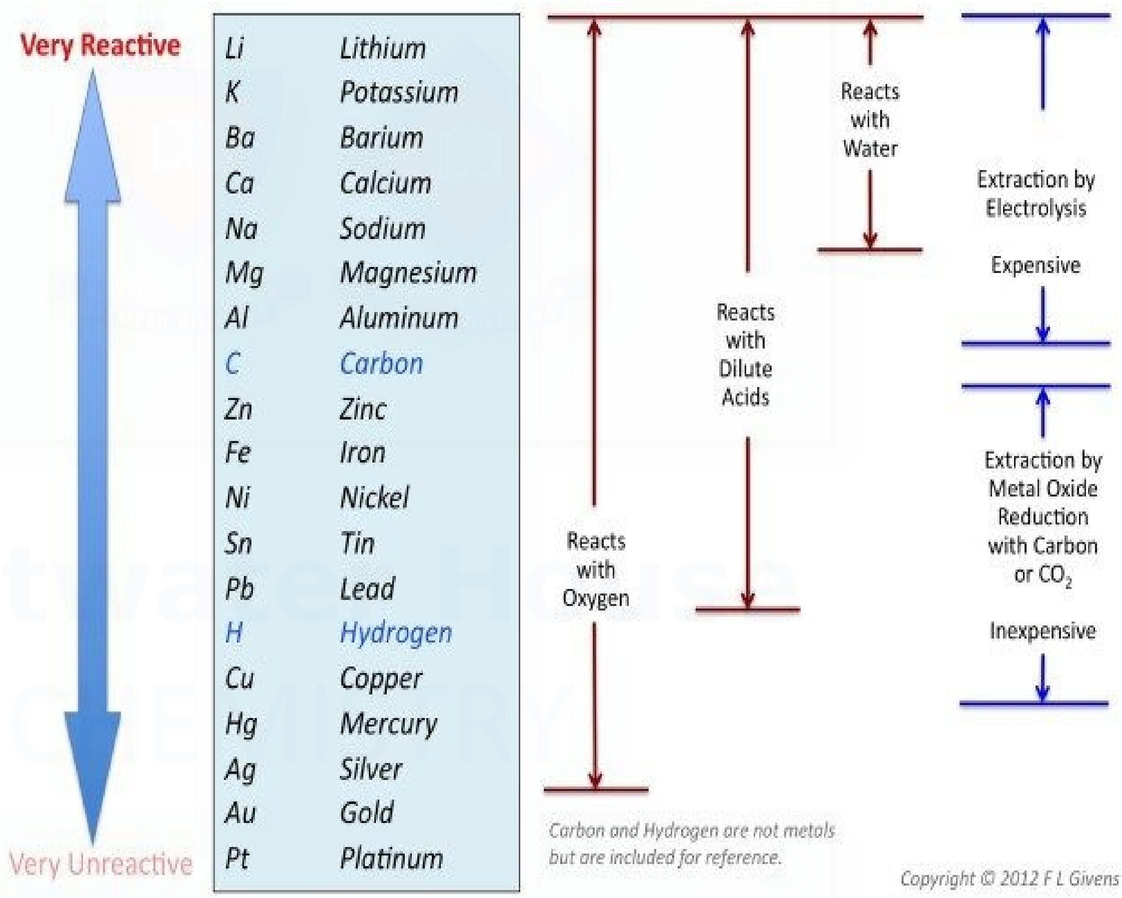
Reaction with oxygen, water and dilute acid.

Only very reactive metals will react with water.

Less reactive metals will react with dilute acid.

Most metals will react with oxygen only.

Within each reactant, classify which is more reactive based mostly off qualitative observations.



Other metal reactions to determine reactivity series

Metal + water → metal hydroxide + hydrogen
Sodium + water → sodium hydroxide + hydrogen
(balanced chem equation) $2Na_{(s)} + 2H_2O_{(l)} \rightarrow NaOH_{(aq)} + H_{2(g)}$
Metal + oxygen → metal oxide
Aluminium + oxygen → aluminium oxide
(balanced chem equation) $4Al_{(s)} + 3O_{2(g)} \rightarrow 2Al_2O_{3(s)}$

Metal DIsplacement

- A metal is more reactive than another IF it will displace another metal from solution.
- A solution of metal 1 M_1(aq) exists and metal 2 M_2(s) is placed in the solution. If metal 1 becomes solid and metal 2 becomes aqueous, metal 2 has displaced metal 1. Therefore metal 2 is more reactive.

- Copper sulfate solution - $CuSO_4(aq)$ solution exists.
- Magnesium metal is placed in. Magnesium starts to dissolve an dcopper starts to become solid, magnesium has displaced copper from solution → magnesium is more reactive than copper because it displaces copper from solution.
- $CuSO_4(aq) + Mg(s) \rightarrow MgSO_4(aq) + Cu(s)$

Metal displacement reactions - ½ equations

- Take an example of a metal displacement reactions

(1) $Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$
- Full ionic: $Fe_{(s)} + Cu^{2+}_{(aq)} + SO_4{}^{2-}{}_{(aq)} \rightarrow Fe^{2+}_{(aq)} + SO_4{}^{2-}{}_{(aq)} + Cu_{(s)}$
- Net ionic: $Fe_{(s)} + Cu^{2+}_{(aq)} \rightarrow Fe^{2+}_{(aq)} + Cu_{(s)}$

(2) $Mg_{(s)} + FeSO_{4(aq)} \rightarrow MgSO_{4(aq)} + Fe_{(s)}$
- Full ionic: $Mg_{(s)} + Fe^{2+}_{(aq)} + SO_4{}^{2-}{}_{(aq)} \rightarrow Mg^{2+}_{(aq)} + SO_4{}^{2-}{}_{(aq)} + Fe_{(s)}$
- Net ionic: $Mg_{(s)} + Fe^{2+}_{(aq)} + \rightarrow Mg^{2+}_{(aq)} + Fe_{(s)}$

(3) $Zn_{(s)} + 2AgNO_{3(aq)} \rightarrow Zn(NO_3)_{2(aq)} + 2Ag_{(s)}$
- Full ionic: $Zn_{(s)} + 2Ag^{+}_{(aq)} + 2NO_3{}^{-}{}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + 2NO_3{}^{-}{}_{(aq)} + 2Ag_{(s)}$
- Net ionic: $Zn_{(s)} + 2Ag^{+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + 2Ag_{(s)}$

<u>½ equations</u>

- ½ equations help us isolate electron movement
- They usually only involve one element
- When electrons are released (a) we call this oxidation
- When electrons are captured (b) we call this reduction
- Redox → reduction / oxidation. **OIL RIG - oxidation is loss, reduction is gain**
- Write the oxidation / reduction equations for (2) and (3)
    - (2) $Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + 2e^-$ oxidation
      $Fe^{2+}_{(aq)} + 2e^- \rightarrow Fe_{(s)}$ reduction
    - (3) $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$ oxidation
      $2Ag^{+}_{(aq)} + 2e^- \rightarrow 2Ag_{(s)}$ reduction

Terms	How does this relate to oxidization
Ionization energy	Both processes involve the loss of electrons
Atomic radius	Large atomic radius will likely undergo oxidation easily – less electrostatic force holding electrons to nucleus
Electronegativity	Linked: Both processes where electrons are attracted to the nucleus

Metal in Reactive order (Most to least)	Ionization energy (kj/mol)		Atomic Radius (A) (pm)	Electronegativity
Potassium (1^+)	419	3051	280	0.82
Sodium (1^+)	496	4562	227	0.93
Magnesium	738	1451	173	1.31
Aluminum	578	1817	143	1.61
Copper	745	1958	128	1.9
Silver	731	2073	172	1.93
Calcium	590	1145	231	1
Zinc	906	1733	139	1.65

Oxidation or Reduction

Oxidation is loss of electrons
Reduction is gain of electrons

$Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^-$	Oxidation
$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$	Reduction
$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$	Oxidation
$Ag_{(s)} \rightarrow Ag^+_{(aq)} + e^-$	Oxidation
$Mg^{2+}_{(aq)} + 2e^- \rightarrow Mg_{(s)}$	Reduction
$Au_{(s)} \rightarrow Au^{2+}_{(aq)} + 2e^-$	Oxidation
$K^+_{(aq)} + e^- \rightarrow K_{(s)}$	Reduction
$Sn_{(s)} \rightarrow Sn^{2+}_{(aq)} + 2e^-$	Oxidation
$Fe^{3+}_{(aq)} + e^- \rightarrow Fe^{2+}_{(aq)}$	Reduction
$Al_{(s)} \rightarrow Al^{3+}_{(aq)} + 3e^-$	Oxidation

Elemental species are zero 0
Oxidation state of Oxygen in O_2
Oxygen gas is an elemental species. Therefore the oxygen atom has an oxidation state of 0
Oxygen is always -2 (except in peroxides [O-O] bonds where it is -1)

Reductant summary:
Reactive solid metals = strong reductants (will throw away electrons)
Reactive aqueous metals = very weak reductants (no more available electrons to throw away)
Unreactive solid metals = weak reductants (hold onto electrons tightly)
Underactive aqueous metals = good reductant (wants electrons)

	Reductant?	Why?	Oxidant?	Why?
Reactive solid metal	YES	Wants to give away it	No	Has a tendency to loose not gain electrons
Reactive aqueous metals	No	Already given away its electrons, no more to give away	No	Takes lots of energy to make an aqueous reactive metal a solid
Unreactive solid metal	No	Holds onto electrons tightly	No	Does not want any more electrons - stable metal atom
Unreactive aqueous metal	No	Wants to become solid and steal electrons from something	Yes	Wants to get back to solid state and grab electrons from almost anywhere

Spontaneity of redox based on reduction potential

- When doing the calculations for voltage notice we are doing this operation
- Oxidation voltage - reduction voltage = +ve voltage
- **Positive voltage means this is a spontaneous reaction**
- If we do this
- Oxidation voltage - reduction voltage = -ve voltage
- **Negative voltage means this is not spontaneous**