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This book is dedicated to my mum and dad who instilled in me a love of science, my wife, family and grandchildren who encouraged me to persevere with the endeavour and to Ted who made it necessary.

Advice from the Inspector

The notes in this section were taken at a lecture given to chemistry teachers by the chief chemistry inspector at a chemical education conference in University College Cork. In this lecture he told us the advice he would give to students at the start of their chemistry course. I think this is advice well worth taking and it is summarised below.

- **Course coverage is absolutely essential.** You cannot afford to leave out any major part of the course, especially organic chemistry.
- Don't take fifth year as a rest period, this is when you can build up a solid base for the final run in to the exam, and it also means that if you do this you will not be nearly as stressed as the exam approaches which makes learning easier.
- Develop an **understanding** of the subject. Don't just learn it off, because the **questions are phrased in such a way as to test your understanding**, not just your ability to memorise. However memorising is important.
- **Engage with the practical work.** This means learn and understand the experiments and demonstrations specified in the syllabus. There is a minimum of three questions on practical work and two are compulsory, so it is worth a minimum of 25%.
- **Don't leave out organic chemistry** as there is a practical question plus two to two and a half other questions on this topic in the exam. Thus it is a very major part of the course.
- **Develop clarity and accuracy in your answering.** You do this by practice using the marking schemes.*

- **Study past papers and their marking schemes.** These marking schemes have to contain all the possible answers and can appear daunting. You only need one possible answer so cross out the answers that do not apply to you and what you have covered in class, in this way the answers become very brief, and less daunting.
- **Read what you write.** You would be amazed at the strange things that pupils can write under exam pressure, whole words or phrases can be left out or put in the wrong place. If you read over what you have written, especially near the end of the exam, you will avoid this type of error.
- Get on with it.
- Past papers and marking schemes are available at celticpress.ie or examinations.ie

The Exam

- **You must do two out of the first three questions** which are based on the mandatory practical activities and demonstrations. You can do all three and this can be a good strategy. Practical work makes up a minimum of 25% of your exam. This can rise to 33% or more as the other questions also contain practical work.
- Under normal circumstances the three practical questions follow a basic pattern: one question deals with titration, one with organic chemistry and one with the remaining experiments.
- Question four is made up of 11 short questions and you are required to answer eight parts. **You should attempt all parts, as the best eight answers will be taken.** What this means is that you can answer some parts of the question wrongly and still score full marks.
- In question four topics such as the following come up regularly. Famous scientists; shapes of molecules; number of protons, neutrons and electrons in atoms or ions; s,p,d,f configuration of atoms and ions; conjugate acid/base pairs; identifying anions or cations; definitions; name or give the formula of an organic compound. Make sure you can do all of these.
- You must answer a total of eight questions.
- If you have answered your eight questions go back over the questions and your answers carefully making sure you have not missed out any parts of the question and that you have answered the question asked.
- If having done this you still have time left you should attempt to answer another question. Remember your best eight answers will be taken [but two of the first three questions must be included].
- Equilibrium is a good topic if you can do it because it can be answered very quickly.
- pH calculation is often worth half a question and does not involve a great deal of learning.
- Atomic structure and periodic table is usually worth more than one question.
- Organic chemistry is worth about three questions under normal circumstances.

Legend

Higher level Only: Indicated by a coloured bar at the right-hand edge of page. →

Definitions to be learned verbatim:

Example of Definition

Worked Examples

WORKED EXAMPLE

A sample of chlorine was found to consist of 75% chlorine 35 and 25% of chlorine 37. Calculate the relative molecular mass of chlorine.

Experiments

EXPERIMENT: Identifying Anions [negative ions]

- **Sulphate** SO_4^{2-} , add Barium chloride / white precipitate / of barium sulphate / doesn't dissolve in dil. HCl

a

Chapter Sections are indicated throughout by a boxed letter in left-hand margin. Use these as Revision Sections.

Important Information is in bold type.

ATOMIC STRUCTURE – PERIODIC TABLE

a 1.1 Periodic Table

The periodic table is a list of elements arranged so as to demonstrate trends in their physical and chemical properties. It has developed over the centuries as we have gained more knowledge. – A fully functional table is at the end of the booklet.

You do not have to learn any of the numbers: you will be given them in the mathematical table book which contains a Periodic Table, a Table of Electronegativities and a Table of First Ionisation Energies and lots of other useful information. Make sure you are familiar with it and know how to use it before you go into the exam.

- Bring a reliable calculator with new batteries and make sure you know how to use it e.g. to work out pH values using logs.

The Periodic Table of Elements

Electronegativity increases →

Atomic Radius decreases →

S Block P Block

1 2 3 4 3 2 1 0

I II III IV V VI VII 0

VALENCY Group

METALS NON-METALS

D Block

Transition Elements

Electronegativity decreases ↓

Atomic Radius increases ↓

F Block

Lanthanoid Series

Actinoid Series

KEY

☐ = Solid at room temperature ☁ = Gas at room temperature ⚡ = Artificially Made

💧 = Liquid at room temperature ☢ = Radioactive

If the elements of the periodic table are arranged by order of atomic weight there are some differences e.g. K and Ar would change places and be totally out of position.

b 1-2 Groups

Alkali Metals – Group I

- Soft enough to cut easily with a knife and become softer down the group
- More reactive down the group
- All have one electron in their outer shell
- Form ions with 1+ charge
- React with cold water producing hydrogen
- $M + H_2O = MOH + \frac{1}{2} H_2$
- **Stored under oil to prevent reaction with air and water**

Halogens - Group VII

- Become less reactive down the group
- Colour intensifies down the group
- Gas – liquid – solid down the group
- All have seven electrons in their outer shell
- Form ionic bonds with groups I and II
- Form covalent bonds with groups III, IV, V, VI and VII

Alkaline Earth Metals – Group II

- Harder than group 1
- Less reactive than group 1
- More reactive down the group
- All have two electrons in their outer shell
- Form ions with 2+ charges
- React with steam producing hydrogen
- $M + H_2O = MO + H_2$

Noble Gases – Group 0

- All are gases
- All are monatomic [composed of one atom]
- All are unreactive because they have a full outer shell

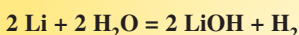
c EXPERIMENT: Demonstration Reaction of Alkali Metals with Water

Lithium in Water

Place some cold water in a beaker and add some phenolphthalein indicator.
Add some freshly cut lithium and note what happens.
It floats on the surface.

It bubbles vigorously but does not ignite.

The phenolphthalein turns purple showing that an alkali has been formed.



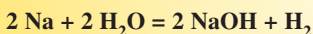
Sodium in Water

Place some cold water in a beaker and add some phenolphthalein indicator.
Add some freshly cut sodium and note what happens
It floats on the surface and will melt

It reacts much more vigorously than lithium and the hydrogen may ignite

It may burn with a yellow flame

The phenolphthalein turns purple showing that an alkali has been formed



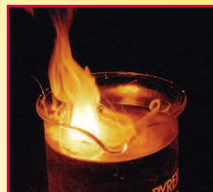
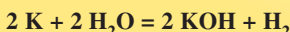
Potassium in Water

Place some cold water in a beaker and add some phenolphthalein indicator.
Add some freshly cut Potassium and note what happens
It floats on the surface and melts

It is so reactive that it almost always bursts into flames

The flames are lilac / purple

The phenolphthalein turns purple showing that an alkali has been formed



ATOMIC STRUCTURE – PERIODIC TABLE

d 1.3 Development of the Periodic Table

Below is a list of famous scientists who made major contributions to the development of the Periodic Table. They are listed in chronological order of their contributions.



Boyle Lavoisier Davy Dalton Dobereiner Newlands Mendeleev Mosley

Boyle (Robert)

Elements can't be broken down into anything simpler
Boyles Law – volume is inversely proportional to pressure at constant temperature

Lavoisier (Anton)

Listed elements known at the time
Law of Conservation of Mass – matter is neither created nor destroyed in a chemical reaction, simply rearranged

Davy (Humphrey)

Discovered New Elements Na, K, Ca etc. using Electrolysis

Dalton (John)

Atomic Theory - element made of tiny particles – all particles of element identical – compounds are chemical combinations of these particles

Dobereiner (Johann)

Triads – groups of 3 elements with similar properties – middle one intermediate in properties.

Two triads are shown below.

Element	Mass	Hardness	Element	Mass	Colour
Lithium	7	Hard	Chlorine	35.5	Green
Sodium	23	Soft	Bromine	80	Red
Potassium	39	Very soft	Iodine	127	Purple

Newlands (John)

Law of Octaves – properties repeat every eighth element
 when arranged in order of mass – no noble gases because they had not been discovered at that time.

Mendeleev (Dmitri)

Periodic Table – Arranged elements by property space rather than just weight – left gaps – predicted properties of elements to go in gaps – e.g. Scandium

Moseley (Henry)

Gave us **Atomic Number** – i.e. the number of protons in an atom

e Modern Table v Mendeleev's

Modern Table	Mendeleev's Table
NO Gaps	Gaps left for undiscovered elements
Noble gases present	Noble gases absent
Transition elements in a separate block	Transition elements NOT in a separate block
Elements in order of atomic number	Elements in order of atomic mass

Symbols

You must know the **symbols of first 36 elements**

Groups

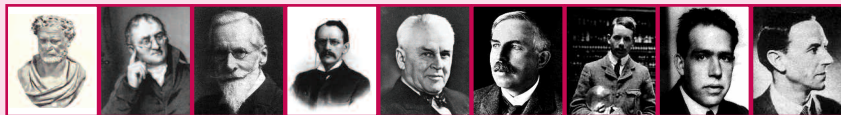
These are the vertical columns in the Periodic Table. Some have got specific names see below.

Group Number	Name
I	Alkali Metals
II	Alkaline Earth Metals
VII	Halogens
O	Noble Gases

Periods

These are the horizontal rows in Periodic Table. The outer shell of the atom is filling one electron at a time, as one goes across the table from left to right.

f 1.4 Atomic Structure – History of its Development



Democritus Dalton Crookes Thompson Millikan Rutherford Mosley Bohr Chadwick

Democritus

Ancient Greek gave us idea of **atoms as indivisible** plus 4 elements earth, wind, fire and water

Dalton (John)

Atomic Theory - element made of tiny particles – all particles of element identical – compounds combinations of these particles

Crookes (William)



Maltese Cross Experiment –

Hot negative electrode [cathode] produced rays that caused a phosphorescent screen to glow in a **vacuum tube**.

Crookes called these **Cathode Rays**

Rays could be blocked by a metal cross shape and a shadow of the cross is cast on the end of the tube.

Thompson (J J)



Deflected cathode rays with charged plates so they must be **negative**
Called them electrons [name coined by **Johnstone Stoney** of UCG]

Magnet also deflects rays.

Plum Pudding Model of the Atom – negative electrons in a positive dough

Worked out **Charge/Mass Ratio** of the electron

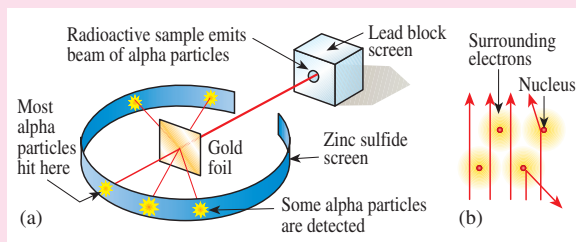
Millikan (Robert)

Discovered the **Mass of Electron**

Oil Drop Experiment – charged oil drops using X-rays and measured charge to stop them falling. Then calculated the mass of the electron using the charge/mass ratio.

Rutherford (Ernest)

Discovered the Nucleus using Gold Foil Experiment -



- **Fired alpha particles / from radioactive source / in block of lead / at very thin gold-leaf**
- Alpha particles seen as flashes of light on fluorescent screen

- **Almost all passed through** – so atom mostly empty space – nucleus very small
- **Some deflected** - so nucleus positive – because like charges repel
- **Some bounced straight back** - so nucleus is very dense [has a lot of mass]

Moseley (Henry)

Each element has a characteristic positive charge. He called this the **Atomic Number**

Bohr (Niels)

Element is a substance whose atoms all have same atomic number.

Electrons orbit nucleus in shells – energy quantised i.e. only certain values allowed

Maximum number of electrons in shell = $2n^2$
[n = shell number]

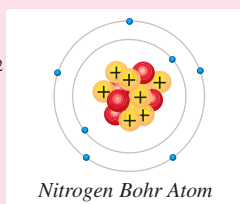
Rutherford (Ernest)

Discovered **protons** when bombarding nuclei with alpha particles.

Suspected other particles were needed to hold protons together

Chadwick (James)

Discovered Neutrons by bombarding Be with alpha particles



Nitrogen Bohr Atom

ATOMIC STRUCTURE – PERIODIC TABLE

g Matter

Anything that has mass

Matter is made of Particles

Atoms	No charge, tiny – 12g of carbon has 600 thousand million, million, million, atoms [6×10^{23}]
Molecules	Groups of atoms chemically joined
Ions	Atoms or groups of atoms which have a charge

Sub-atomic Particles

Particle	Mass	Charge	Location
Proton	1	+1	Nucleus
Neutron	1	neutral	Nucleus
Electron	0	-1	Orbiting nucleus

Atomic Number (Z) = Number of Protons. This determines what the element is.

Mass Number (A) = Number of Protons plus the Number of Neutrons in an atom or isotope

Sometimes referred to as the **Nucleon Number**

Relative Atomic Mass (A_r)

is the average mass of the element as it occurs in nature compared to $\frac{1}{12}$ of the mass of ^{12}C isotope [when the isotopes and their abundance are taken into account]

Number of Protons = atomic number

Number of Electrons = number of protons = **atomic number** [in atoms]

In **ions** add an electron for each negative charge and subtract an electron for each positive charge

Number of Neutrons = Mass Number – Atomic Number

h Isotopes

Atoms of an element which have the same atomic number but different mass numbers due to different numbers of neutrons

Isotope	Atomic No.	Mass No.	Protons	Neutrons	Electrons	e ⁻ Pattern
Carbon 12	6	12	6	6	6	2,4
Carbon 14	6	14	6	8	6	2,4
Hydrogen 1 <i>Protium</i>	1	1	1	0	1	1,
Hydrogen 2 <i>Deuterium</i>	1	2	1	1	1	1,
Hydrogen 3 <i>Tritium</i>	1	3	1	2	1	1,

Numbers of electrons [and protons] are the same therefore chemical properties are identical. Not all isotopes are radioactive but ^{14}C is

Relative Atomic Mass of Element with Isotopes

$$A_r = \frac{(\text{Mass of Isotope 1} \times \% \text{ abundance}) + (\text{Mass of Isotope 2} \times \% \text{ abundance})}{100}$$

WORKED EXAMPLE

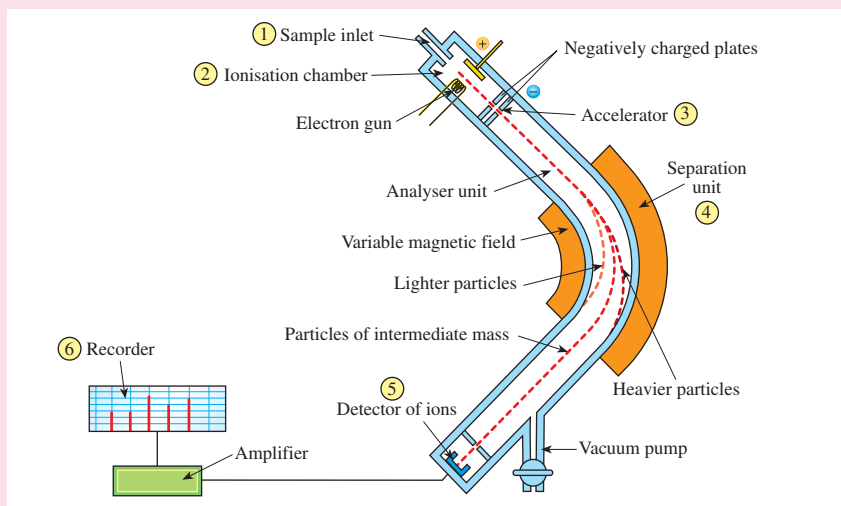
A sample of chlorine was found to consist of 75% chlorine 35 and 25% of chlorine 37. Calculate the relative molecular mass of chlorine.

$$A_r = \frac{(35 \times 75) + (37 \times 25)}{100} = \frac{(2625 + 925)}{100} = \frac{3550}{100} = 35.5$$

1.5 Mass Spectrometer

In works on the principle:-

- ions can be separated
- by mass
- while moving in a magnetic field



It was invented by **F W Aston** to investigate isotopes of Neon

Learn how to label the diagram above

Fundamental Processes in the mass spectrometer

Process	Activity
1. Vaporisation	The substance is vaporised by heat and reduced pressure
2. Ionisation	The sample is ionised by bombarding it with high energy electrons producing positive ions
3. Acceleration	These ions are accelerated by negatively charged plates
4. Separation	These ions are then deflected by a magnetic field as they pass through it. Light particles are deflected more than heavier particles.
5. Detection	The ions are detected by a sensitive plate at the end of the spectrometer
6. Display	Both the mass and abundance of each substance is displayed on a screen

Uses: -

- Determining both the mass and abundance of isotopes
- Analysis of
 - (i) gases from a waste dump
 - (ii) trace organic pollutants in water.

ATOMIC STRUCTURE – PERIODIC TABLE

j 1.6 Radioactivity

The spontaneous breaking up of an unstable nucleus with the emission of one or more types of radiation



Discovered by **Henri Becquerel** while working on uranium salts

Geiger-Muller Tube is used to detect radiation



Becquerel



Curies

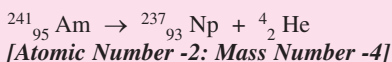
Curies discovered two new elements

Polonium 84 [after Poland] and **Radium 88** [giver of rays]

Radiation Types and Properties

Type	Composition	Penetrating Power	Source	Use
Alpha Particles	2 Protons + 2 Neutrons i.e. a <u>He nucleus</u>	Few cm of air	Americium 241	Smoke alarms
Beta Particles	High energy electrons emitted from nucleus	5 mm Al	Carbon 14	Carbon dating
Gamma Rays	Electromagnetic radiation	Few cm lead	Cobalt 60	Killing tumours

Alpha Particle Emission Equation



Beta Particle Emission Equation



Neutron ejects an electron and becomes a proton – atomic number goes up by one mass stays the same.

Radioisotopes

Not all isotopes are radioactive. Those that are, are called **radioisotopes**

Half -life

This is the time taken for half of the nuclei in any given sample to decay

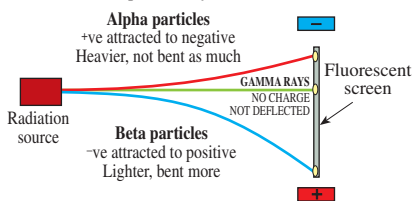
Uses of Radioactivity [3 examples need to be known]

- **Archaeological - Carbon Dating** – ${}^{14}\text{C}$ in living tissue. The ratio of ${}^{12}\text{C}$: ${}^{14}\text{C}$ is constant in living tissue. After death ${}^{14}\text{C}$ decays and the proportion of ${}^{14}\text{C}$ drops in proportion to time elapsed.
- **Medical – killing tumours** and sterilising surgical equipment - ${}^{60}\text{Co}$
- **Food Irradiation – increases shelf life** - ${}^{60}\text{Co}$

Radioactivity is everywhere

- **Radon in granite** – causes cancer if inhaled for a long period
- **Your bones contain radioactive carbon 14.**

To show the Properties of Radioactive Emissions



1.7 Electronic Structure of Atoms

Electrons can only be found at certain distances from the nucleus.

The **further the e^- is from the nucleus the more energy it has.**

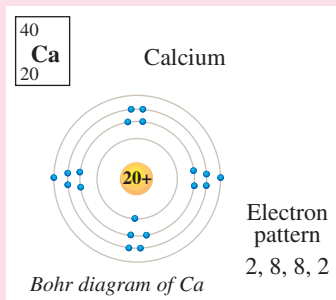
The electron can only have certain energies and these energies are said to be **quantised**.

These main energy levels can contain certain a fixed number of electrons, e.g. in calcium.

$$1^{\text{st}} = 2, 2^{\text{nd}} = 8, 3^{\text{rd}} = 8, 4^{\text{th}} = 2$$

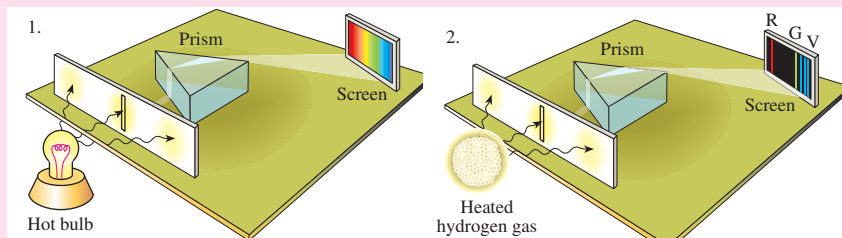
You need to be able to draw Bohr diagrams for first 20 elements.

For the first twenty elements the group number tells us the number of electrons in the outer shell of the elements and this determines its chemical properties. So members of a group have the same number of electrons in their outer shell and thus react in similar ways.



Atomic Structure

We have discovered most of what we know about the electron arrangement in atoms from studying spectra.



There are 3 types of spectra

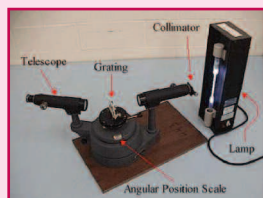
1. Continuous – from sun **2. (Line) Emission** – from heated or shocked elements

3. (Line) Absorption – certain frequencies removed as light passes through gaseous element (see following page)

1.8 Emission Spectra

Metals excited by heat or high voltage electricity give out characteristic colours of light. If this light is viewed after passing it through a diffraction grating, or prism, a series of lines can be seen. This is called a **line spectrum** and is **characteristic of each element**.

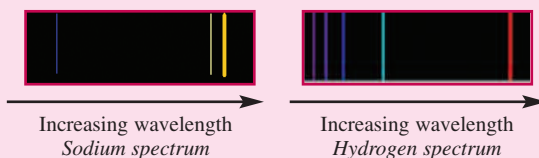
Emission spectra can be produced by flame tests or by special lamps containing the element in question called **discharge tubes**.



Spectroscope with diffraction grating set up in front of a discharge tube lamp

These lines are in the **visible** section of the spectrum and are called the **Balmer Series**.

This is what they look like through a spectroscope.



ATOMIC STRUCTURE – PERIODIC TABLE



Sodium street lamp

Sodium emits most of its light in the orange/yellow region of the spectrum. Orange street lights contain sodium and this explains their colour. Fireworks contain different metals or their compounds and the explosion produces heat which causes them to emit their characteristic colours.



Potassium firework

Emission Line Spectrum Formation

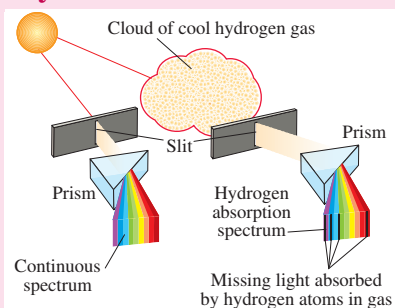
- Electron is normally in its lowest available energy level – **Ground State (E_1)**
- Electron **excited by heat or electricity** [high voltage]
- Jumps to higher energy level – **Excited State (E_2)**
- It is **Unstable**
- Drops back** to lower level
- Energy difference** between the two levels is **released as Photon or Electromagnetic Radiation**
- Frequency** of energy emitted is **proportional to drop** (difference in energy)
- $(E_2 - E_1) = hf$ **h is Planck's Constant** and **f is frequency** of radiation
- Spectral lines formed are evidence of energy levels**
- electrons falling into $n = 1$ give UV radiation [Lyman Series]
- electrons falling into $n = 2$ give **visible light** [Balmer Series] *You can be asked this one*
- electrons falling into $n = 3, 4, 5$ give IR [Paschen, Brackett, Pfund Series]

m 1.9 Atomic Absorption Spectrometry

When light passes through an element in its gaseous phase it **absorbs** wavelengths of light identical to those it emits when it is excited. These appear as dark lines in the spectrum and it is called an **absorption spectrum**.

The amount of an element present in a sample can be calculated using an **atomic absorption spectrometer**.

The sample is converted to atoms in a flame. Light of a particular wavelength which is characteristic to the element being looked for is passed through it. The amount of this light absorbed is proportional to the abundance of the element in the sample. (see p. 81)



Use: determining the presence and concentration of heavy metals e.g. lead in a blood or water sample

m 1.10 Electron Configuration

Energy Level

Shell which electrons of equal energy can occupy

n is the **main energy level** [shell] or **Principle Quantum Number**.

Number of electrons that a main energy level can hold is $2n^2$

It was later found that the main energy levels were divided into **sublevels**

There are up to 4 of these **s, p, d and f** in any main energy level

Main Energy Level	Sublevels present
$n = 1$	1s
$n = 2$	2s < 2p
$n = 3$	3s < 3p < 3d
$n = 4$	4s < 4p < 4d < 4f
$n = 5$	5s < 5p < 5d < 5f

Sublevels have Different Energies

Within a main energy level the sublevels are arranged from lowest energy to highest energy

$s < p < d < f$

Heisenberg's Uncertainty Principle

The more accurately one determines the position of an electron the less accurately one can determine its velocity.

Electrons have Wave-particulate Duality

Electrons have properties of both waves and particles at same time

Sublevels are further divided into **orbitals of equal energy**

Orbitals

Region in space around the nucleus of an atom in which electrons are most likely to be found.

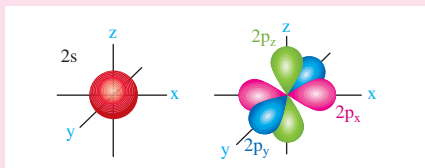
Within a sublevel the **orbitals all have equal energy** e.g. $p_x = p_y = p_z$

Each **orbital can hold 2 electrons [with opposite spins]**

Sublevel type	Orbitals present	Shape	Maximum number of e^-
s	1 (s)	spherical	2
p	3 (p_x, p_y, p_z)	dumbbell - mutually at right angles	6
d	5 - - - -	complex - don't need to know	10
f	7 - - - -	complex - don't need to know	14

Orbital Shapes

You need to be able to draw these individually or as a group in the case of p orbitals



Electron Configuration [s, p, d, f. pattern]

Aufbau Principle

Electrons occupy the lowest available energy level

Order of Filling

Within a sublevel the orbitals are arranged from lowest energy to highest energy

$s \rightarrow p \rightarrow d \rightarrow f$

$1s$
 $2s, 2p$
 $3s, 3p, 3d$
 $4s, 4p, 4d, 4f$
 $5s, 5p, 5d, 5f$

The orbitals are filled from the lowest energy to highest energy.

$1s \rightarrow 2s \rightarrow 2p \rightarrow 3s \rightarrow 3p \rightarrow 4s \rightarrow 3d \rightarrow 4p \rightarrow 5s \rightarrow \text{etc.}$

The diagram to the left is easier to remember than this sequence.

Note that **4s** is lower in energy than **3d** and so is filled before **3d**.

Number of electrons [in an atom] = atomic number

You need to learn how to produce the electron patterns for first 36 elements

WORKED EXAMPLE

Chlorine

Atomic number 17 therefore 17 electrons

$1s^2, 2s^2, 2p^6, 3s^2, 3p^5$

Vanadium

Atomic number 23 therefore 23 electrons

$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^3$

ATOMIC STRUCTURE – PERIODIC TABLE

0 Exceptions

WORKED EXAMPLE

Chromium $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1 3d^5$

Half full s and d sublevel is more stable than full s and $3d^4$.

Copper $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1 3d^{10}$

Half full s and full d sublevel is more stable than full s and nearly full d orbital

Electron Configuration of Ions

Each plus charge = 1 electron lost.

Each minus charge = 1 electron gained

WORKED EXAMPLE

Calcium ion Ca^{2+}

2+ charge thus it has lost 2 electrons therefore $[1s^2, 2s^2, 2p^6, 3s^2, 3p^6]^{2+}$

Phosphide ion P^{3-}

3- charge thus it has gained 3 electrons therefore $[1s^2, 2s^2, 2p^6, 3s^2, 3p^6]^{3-}$

When writing the pattern of an ion put the pattern in square brackets with charge outside

N.B. [both have pattern of Ar – the nearest Noble gas - it is their charge that makes them different]

Only patterns of ions of first 20 elements required

Arrangement of Electrons in Orbitals of Equal Energy

See arrangement in second period elements to the right

Hund's Rule of Maximum Multiplicity

When two or more orbitals of equal energy are available, the electrons occupy them singly and then in pairs.

Pauli Exclusion Principle

No more than two electrons may occupy an orbital and to do so they must have opposite spin.

Arrows are in opposite directions because the electrons spin in opposite directions

Each electron can be specified by 4 numbers -

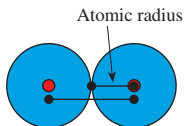
Principal Quantum Number - Sublevel - Orbital - Spin

Electron arrangement in 2p orbitals

Boron $2p^1$	\uparrow			
Carbon $2p^2$	\uparrow	\uparrow		
Nitrogen $2p^3$	\uparrow	\uparrow	\uparrow	Stable
Oxygen $2p^4$	$\uparrow\downarrow$	\uparrow	\uparrow	
Fluorine $2p^5$	$\uparrow\downarrow$	$\uparrow\downarrow$	\uparrow	
Neon $2p^6$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	Very stable

P 1.11 Atomic Radius

Half the distance between the nuclei of two adjacent atoms of the same element joined by a single covalent bond



Trends in Atomic Radius

Across Table	<ul style="list-style-type: none"> decreases due to greater nuclear charge but same shielding effect
Down table	<ul style="list-style-type: none"> increases due to new shell being added so more shielding increased nuclear charge - not as significant
Noble Gases	<ul style="list-style-type: none"> don't bond so they have NO atomic radius

q 1.12 Electronegativity

Measure of the attractiveness of an atom for a shared pair of electrons in a covalent bond.

Invented by **Linus Pauling**

Difference in electronegativity determines the type of bonding between atoms

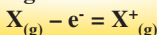
Difference	Bond type
0 to 0.4	Pure covalent
0.4 to 1.7	Polar covalent
> 1.7	Ionic

Trends in Electronegativity in the Periodic Table

Across	Increases – bigger nuclear charge – smaller atomic radius
Down	Decreases – further from nucleus – and more shielding – this is more important than increased nuclear charge
Noble Gases	do not have electronegativities - as they do not form bonds

r 1.13 First Ionisation Energy

Minimum energy required to completely remove the most loosely held electron from 1 mole of gaseous atoms in their ground state



Don't forget to put in the state symbols.

Trends in First Ionisation Energy

Across Table	<ul style="list-style-type: none"> Value goes up due to bigger nuclear charge and smaller atomic radius [i.e. electron nearer nucleus.]
Down Table	<ul style="list-style-type: none"> value goes down bigger nuclear charge but e⁻ further from nucleus so more shielding / therefore easier to remove electron

Value Higher than Expected at Various Places –

Group II elements due to **filled s sub-level** being more stable

Group V elements due to **half filled p sub-level** being more stable

Group 0 elements **filled p sublevel** being very stable

Some Examples from Graph Opposite

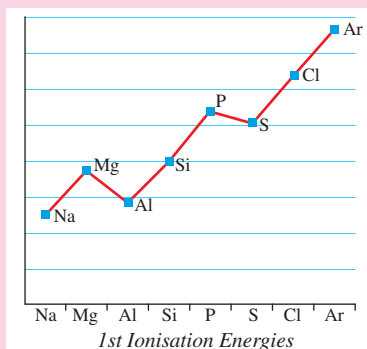
- Mg higher than general trend due to full 3s sublevel
- P is higher due to a half- filled 3p sub-level
- Ar is higher due to full 3p sublevel

Reactivity Trends in Groups I & VII

Group I become **more reactive** as one goes **down** because radius increases and there is more screening which masks the increased nuclear charge. It is easier to lose an electron, and thus react.

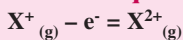
Group VII become **less reactive** as one goes

down because radius increases and there is more screening which makes it more difficult for the nucleus to attract electrons, and thus react.



ATOMIC STRUCTURE – PERIODIC TABLE

Second and Subsequent Ionisation Energies



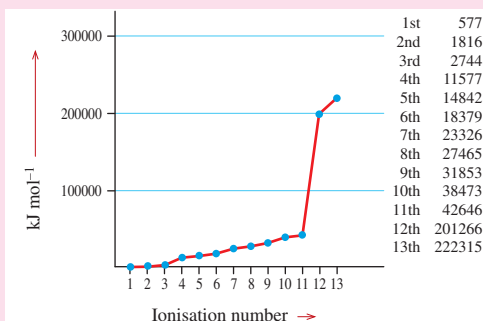
Second higher than the first because removing e^- from

- an ion so there is more positive charges per electron i.e. greater effective nuclear charge.
- closer to nucleus as atomic radius of ion is smaller than corresponding atom.

Bigger increase if an electron is removed from a half filled sub-level

Much bigger if new shell entered because electron being removed is

- closer to nucleus
- in a full sublevel
- has less shielding



The 13 Ionisation Energies of Aluminium

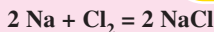
Successive Ionisation Energies of Aluminium Explained

- First three ionisation energies increase steadily as electrons are removed from the third shell
- 4th is a big jump as the second shell is entered
- 5th to 11th get steadily bigger as successive electrons are removed from the second shell
- 12th very big jump as 1st shell is entered

These jumps in ionisation energies are evidence for the existence of energy levels (shells)

S 1.14 Oxidation and Reduction

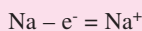
Oxidation is loss of electrons (OIL)
Reduction is gain of electrons (RIG)



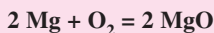
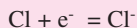
Na loses an electron to become Na^+

Cl gains an electron to become Cl^-

- it is oxidised



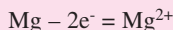
- it is reduced



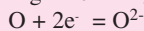
Mg loses 2 electrons to become Mg^{2+}

Oxygen gains 2 electrons to become O^{2-}

- it is oxidised

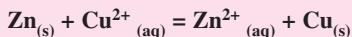


- it is reduced



Displacement Reactions

When Zn is placed in copper (II) sulphate solution the following reaction takes place



Zinc loses 2 electrons to copper – it is oxidised $\text{Zn} - 2e^- = \text{Zn}^{2+}$

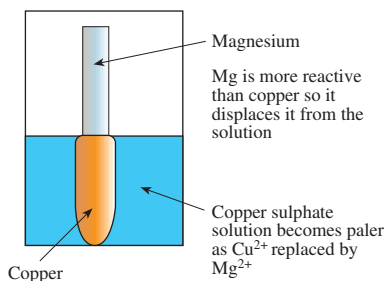
Copper gains 2 electrons from zinc – it is reduced $\text{Cu}^{2+} + 2e^- = \text{Cu}$

This is called a **displacement reaction** as the copper is displaced (kicked out) from the solution

The **more reactive metal displaces the less reactive metal from the solution.**

The more reactive metal is oxidised and the less reactive is reduced.

Uses: to place elements in order of reactivity and extract valuable copper from solutions of its salts using scrap iron.



Redox Reactions

The above reactions involve both **reduction** and **oxidation** and are called **redox reactions**

Rusting of Iron is an example of oxidation

Iron loses 3 electrons when it reacts with water to form Fe^{3+} $\text{Fe} - 3\text{e}^- = \text{Fe}^{3+}$

Oxidising Agent

A substance that allows oxidation to happen by gaining electrons

In the above redox reaction examples chlorine, oxygen and Cu^{2+} are **oxidising agents**

Other examples of common oxidising agents are MnO_4^- and $\text{Cr}_2\text{O}_7^{2-}$

Chlorine is used in swimming pools to kill bacteria by oxidation

Reducing Agent

A substance that allows reduction to happen by losing electrons

Hydrogen, carbon and carbon monoxide are common reducing agents.

Remember the oxidising agent is reduced as it oxidises and the reducing agent is oxidised as it reduces

1.15 Electrochemical Series

A series of metals arranged in order of their ability to be oxidised

Use mnemonics to learn off lists

King Neptune Caught
Many Angry Zulus
Fighting Some Pretty
Horrible Country Afgans

First letter of word matches first
letter of element symbol

Positive ions	Negative ions
K^+ Na^+ Ca^{2+} Mg^{2+} Al^{3+} Zn^{2+} Fe^{2+} Sn^{2+} Pb^{2+} H^+ Cu^{2+} Ag^+	F^- SO_4^{2-} NO_3^- Cl^- Br^- I^- OH^-

For Seven Nights
Charlie Brown
Ingested Hydroxides

1.16 Electrolysis

Chemical reaction caused by the passage of an electric current through a liquid known as the electrolyte

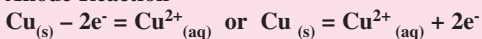
- Electrolyte** - liquid in which electrolysis takes place. Usually an ionic solution but it can also be a fused [melted] ionic compound
- Anode** - positive electrode. Positive because the battery sucks electrons out of it
- Cathode** - Negative electrode. Negative because the battery pumps electrons into it.
Remember which is which using **Cat NAP**, **Catode Negative Anode Positive**
- Anion** - negative ion. Called anion because it is attracted to the opposite charge of the anode
- Cation** - positive ion. Called cation because it is attracted to the opposite charge of the cathode
- Inert Electrodes** - do not react with the electrolyte e.g. graphite and platinum
- Active electrodes** - react with electrolyte e.g. copper and iron

ATOMIC STRUCTURE – PERIODIC TABLE

W Electrolysis of Copper Sulphate using Copper Electrodes

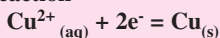
Purifying copper by electrolysis

• Anode Reaction

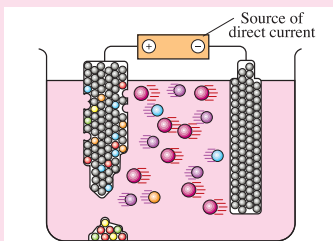


- Anode loses mass as copper dissolves off it
- Impurities [Au, Ag, Pt etc.] fall to bottom

• Cathode Reaction

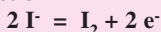


- Cathode gains mass as Cu is deposited on it
- Deposited Cu is 99.9% pure



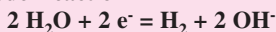
Electrolysis of Potassium Iodide solution using Inert Electrodes

• Anode Reaction

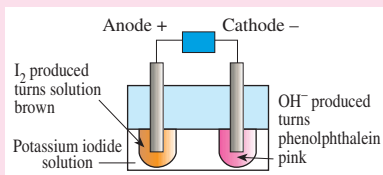


- Solution goes brown as Iodine forms

• Cathode Reaction

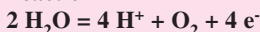


- Phenolphthalein goes pink due to OH^-



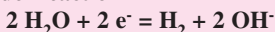
Electrolysis of Sodium Sulphate solution using Inert Electrodes

• Anode Reaction

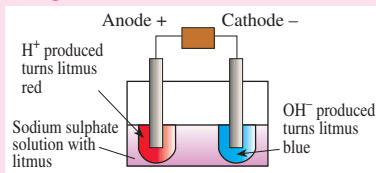


- Litmus goes red due to H^+

• Cathode Reaction



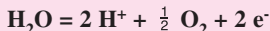
- Litmus goes blue due to OH^-



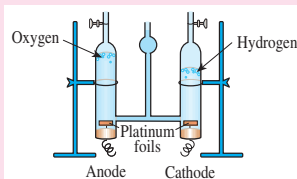
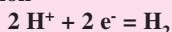
[Multiply by 2 if you want to balance electron movement]

Electrolysis of Acidified Water in the Hofmann Voltameter

• Anode Reaction

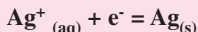


• Cathode Reaction



Electroplating

Covering cathode in metal e.g. covering a copper coin with silver. The copper coin is made into the cathode and placed in a silver nitrate solution



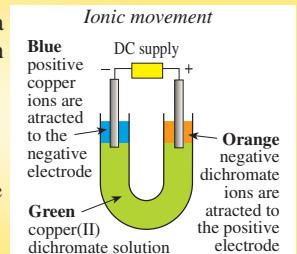
Uses of Electroplating

1. Steel can be electroplated with chromium to make it attractive and corrosion resistant e.g. car bumpers
2. Nickel cutlery can be plated with silver. This is called EPNS electroplated nickel silver.

X 1.17 Experiments

EXPERIMENT: Demonstration Ionic Movement

- Place some green copper (II) dichromate solution in a U-tube Carefully top up both sides of the U-tube with dilute hydrochloric acid. Place carbon electrodes in both sides and switch on the current. The solution around the cathode becomes blue as copper (II) ions $[\text{Cu}^{2+}]$ congregate attracted by the negative charge. The solution around the anode becomes orange as the dichromate ions $[\text{Cr}_2\text{O}_7^{2-}]$ congregate there attracted by the positive charge.



EXPERIMENT: Flame Test to Identify Metal Ions Present in a Salt

- Clean nichrome wire by dipping in conc. HCl and holding it in the flame
Wet wire again with conc. HCl and pick up a sample of salt of the metal selected
Hold it in the edge of a non-luminous bunsen flame. Note the colour produced



Sodium
Yellow



Potassium
Lilac



Lithium
Crimson



Strontium
Red



Barium
Yellow/Green

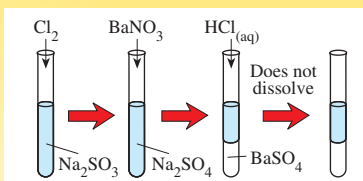


Copper
Blue-Green

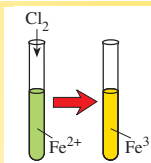
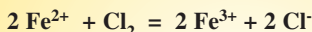
EXPERIMENT: Reactivity of the Halogens

- Add chlorine gas to sodium sulphite (Na_2SO_3) solution and shake. Test by adding barium nitrate (or chloride) - a white precipitate forms.

It does not dissolve when dilute hydrochloric acid is added so the sulphite has been oxidised to a sulphate (Na_2SO_4) by the chlorine.

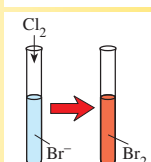
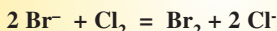


- Add chlorine to pale green iron (II) sulphate solution and mix. It turns brown as iron (II) is oxidised to iron (III) by the chlorine gas.



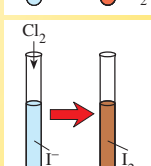
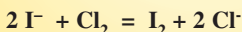
- Add chlorine gas to colourless sodium bromide solution and shake. The solution turns red/brown as the chlorine displaces the bromine from the solution.

This shows that chlorine is more reactive than bromine.



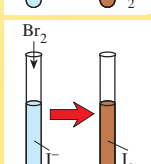
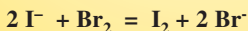
- Add chlorine gas to colourless sodium iodide solution and shake. The solution turns brown as the chlorine displaces the iodine from the solution.

This shows that chlorine is more reactive than iodine.



- Add bromine to colourless potassium iodide solution and shake. It turns brown as bromine displaces iodine from the solution.

This shows that bromine is more reactive than iodine.



These reactions show;

- that the halogens are **powerful oxidising agents**
- that in **order of decreasing reactivity they are chlorine > bromine > iodine**

Questions on this Section found in Past Exams Year by Year

2009	2008	2007	2006	2005	2004	2003	2002
3	4 a,b,c	4 a,f,	4 a,b,c,f	4 a,c,d,e,f	4 a,i,j	3	4 a,b,c,f
4 a,b,c	5	5	5	5a,b	5	4 a,c,,f,i	5
5	10 b,c	11a	10a	10b	10b	5a,b	7c
10 b,c				11a	11a	10c	10b
						11b	11b

CHEMICAL BONDING

a 2.1 Chemical Compounds

Noble Gases are monatomic and unreactive

Stability is due to a full outer shell of electrons

- **Helium** – light and inert so used in airships and divers breathing mixture
- **Argon** – used for filling incandescent light bulbs and as inert atmosphere for welding

Valency

Is the combining power of an atom

i.e. how many atoms it can combine with or how many electrons it has to gain, or lose, or share to attain Noble Gas structure.

Valency	1	2	3	4	3	2	1	0
Group	I	II	III	IV	V	VI	VII	0

Transition Elements (Cu, Fe, Cr and Mn only) show the following properties

- **Variable Valency** [different oxidation states]
- **Coloured ions**
- **Catalytic properties** as do many of their compounds

You have to be able to predict simple chemical formulae of the first 36 elements (excluding d-block elements) of hydroxides, carbonates, nitrates, hydrogencarbonates, sulfites and sulfates of these elements (nitrates, hydrogencarbonates, sulfites and sulfates *higher level only*).

Octet Rule

Having 8 electrons in the outer shell gives an atom or ion stability

Limitations:-

Elements such as H, He, Li, Be, don't follow the rule as they have only two electrons in their outer shell

Al and B have only 6 e^- in outer shell in AlF_3 and BF_3 respectively and these are stable.

b 2.2 Ionic Bonding

Ionic bonding is the result of electron transfer

Each atom is trying to get a full outer shell.

This is achieved by giving and taking of electrons i.e. Total e^- s lost = total e^- s gained

Metals lose electrons to form positive ions (Cations)

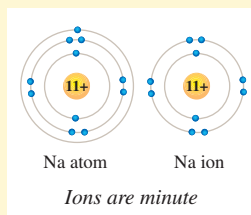
- Group I lose $1e^-$ to form ions with 1+ charge
- Group II lose $2e^-$ to form ions with 2+ charges
- Group III lose $3e^-$ to form ions with 3+ charges

Non-metals gain electrons to form negative ions (Anions)

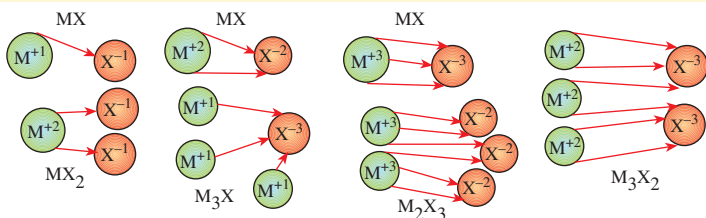
- Group V gain $3e^-$ to form ions with 3- charges
- Group VI gain $2e^-$ to form ions with 2- charges
- Group VII gain $1e^-$ to form ions with 1- charge

Ions are Minute

Positive ions are smaller than the atoms they are derived from



Represent Ionic Bonds Using Dot and Cross Diagrams



Ionic compounds always contain a metal [and a non-metal]

Everyday Uses of Ionic Compounds

- Salt [NaCl] tablets to replace salt lost by sweating
- Iron sulphate tablets to cure anaemia – Fe ions in haemoglobin which carries O_2 .

Characteristics of Ionic Substances

- Regular arrangements of ions forms **Crystal Lattice** and leads to crystals
- **High Melting and Boiling Points** because each ion held in place by many other ions of opposite charge
- **Solids do not conduct** [insulators] because ions are firmly held in place
- **Liquids and solutions conduct with electrolysis** as ions are free to move
- Solids **Brittle**
- Many are **soluble in water** [like dissolves like] but insoluble in organic solvents such as benzene

EXPERIMENT: Identifying Anions [negative ions]

- **Sulphate SO_4^{2-}** , add Barium chloride / white precipitate / of barium sulphate / doesn't dissolve in dil. HCl
- **Sulphite SO_3^{2-}** , add Barium chloride / white precipitate / of barium sulphite / does dissolve in dil. HCl
- **Carbonate CO_3^{2-}** , gives CO_2 when dil. HCl added / CO_2 turns lime water milky / take a fresh sample / gives white precipitate / of Magnesium carbonate / when magnesium sulphate added
- **Hydrogencarbonate HCO_3^{-}** , gives CO_2 when dil. HCl added / CO_2 turns lime water milky / take a fresh sample / does not give white precipitate / when magnesium sulphate added / but does when boiled
- **Nitrate NO_3^{-}** , add Iron(II)sulphate solution / add conc. H_2SO_4 / at angle down test tube / brown ring at interface is positive result
- **Phosphate PO_4^{3-}** , add conc. nitric acid to ammonium molybdate / add mixture to phosphate / heat gently in water bath / yellow precipitate is positive result / redissolves in ammonia solution
- **Chloride Cl^{-}** add acidified silver nitrate solution / white precipitate / of silver chloride / re-dissolves in / ammonia solution

Examination of a Model of the NaCl Crystal – see diagram below.

2.3 Covalent Bonding

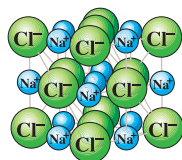
Molecules are tiny but bigger than atoms

Covalent Bonds Involve the sharing of pairs of electrons

Single Bond = sharing **1 pair** of e^{-} (H-H, Cl-Cl)

Double Bond = sharing **2 pairs** of e^{-} ($O=O$)

Triple Bond = sharing **3 pairs** of e^{-} ($N \equiv N$)



*Sodium Chloride
Crystal Structure*

CHEMICAL BONDING

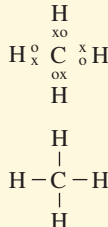
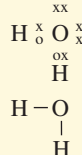
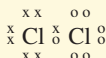
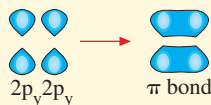
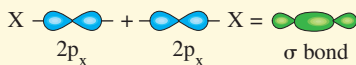
Single Bonds are Sigma Bonds

Double Bonds are a Sigma and a Pi Bond

Triple Bonds are 1 Sigma and 2 Pi Bonds.

Pi bonds are weaker than sigma bonds

Only valence electrons need be shown



Non-Polar Covalent Bonding

– equal distribution of charge between the two atoms in the bond.

Difference in electronegativity between atoms determines the bond type

Electronegativity Difference	0 < 0.4	>0.4 <1.7	>1.7
Bond Type	Covalent	Polar Covalent	Ionic

f

Polarity Test for Liquids

- A charged rod brought towards a thin stream of liquid from a burette
- If polar the stream will be attracted and bends towards rod
- If polarity of rod changed - stream will still bend towards it, but opposite side of molecule will be attracted

Polar (i) Water (ii) Alcohol

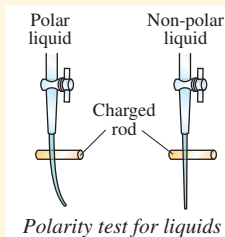
Non Polar (i) Cooking oil (ii) Petrol

Characteristics of Covalent Substances

- Made of individual molecules with characteristic shapes
- Low melting and boiling points
- Usually insoluble in water
- Do not conduct electricity [insulators]

Like dissolves Like

- Polar liquids such as water dissolve ionic compounds
- Non-polar liquids such as methyl benzene and cyclopropane dissolve non-polar substances such as wax.



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2.4 Electronegativity

Measure of the attractiveness of an atom for a shared pair of electrons in a covalent bond

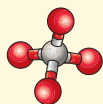
Variations in Electronegativity Caused by:

- Different nuclear charges
- Distance from nucleus
- Screening effect of inner shells

Trends:

– Across Period – Increases	– Down Group – Decreases
Bigger nuclear charge, Smaller atomic radius, Same screening	Bigger Atomic radius [further from nucleus] More screening [by inner electrons] Even though bigger nuclear charge [offset be increase in radius]

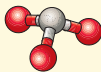
h 2.5 Shapes of Molecules



Tetrahedral



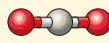
Trigonal Planar



Pyramidal



Angular



Linear

Bond Angles

Bonding pairs are pairs of electrons involved in bonding

Lone pairs are pairs of electrons not involved in bonding

e- pair composition	Shape	Bond angle	Examples
4 Bonding 0 Lone Pairs	Tetrahedral	109.5°	CCl ₄ , CH ₄ , NH ₄ ⁺
3 Bonding 1 Lone Pair	Pyramidal	107°	NH ₃ , H ₃ O ⁺
3 Bonding 0 Lone Pairs	Trigonal [Planar]	120°	FeCl ₃ , AlF ₃ , BF ₃
2 Bonding 2 Lone	Angular [V-shaped]	H ₂ O = 104.5° H ₂ S = 102°	H ₂ O, H ₂ S
2 Bonding 0 Lone	Linear	180°	CO ₂ , C ₂ H ₂

Explain the Difference in Bond Angle between H₂O and H₂S

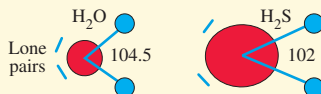
Lone Pairs repel more than Bonding Pairs

Pair Composition	Lone – Lone	Lone – Bonding	Bonding – Bonding
Repulsion	Biggest	Medium	Least

H₂O = 104.5° - 2 lone pairs push 2 bonding pairs even further round [together] than in the tetrahedral arrangement;

H₂S = 102° - In H₂S bonding pairs can get further round due to being further from nucleus before mutually repulsive forces of the bonding pairs have an effect. This reduces the bond angle from 104.5° to 102°

Each Lone Pair reduces the bond angle by 2.5° approx



i 2.6 Intramolecular Bonding [forces within molecules]

Covalent Bonding e.g. diamond – very hard, high melting and boiling points, insulator.

Ionic Bonding – Electrons move from one atom to another, producing ions.

Electrostatic forces hold the ions together. E.g. Na⁺Cl⁻

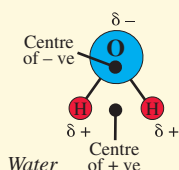
Explain the Relationship between Symmetry and Polarity

Electronegativity – unequal distribution of electrons causes bonds within the molecule to be polar. If centres of + and – charges are not in the same place then the molecule will be polar e.g. water. A molecule with a positive and negative end is called a **dipole**.

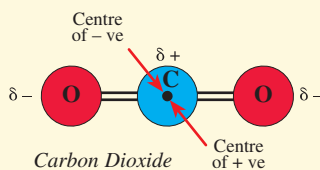
Carbon dioxide is non polar because its centre of + and – are both in the centre of the C atom.

Water is polar because its centre of + and – are in different places.

The centre of + is between the 2 H's while the centre of – is in the O



Water



Carbon Dioxide

CHEMICAL BONDING

j 2.7 Intermolecular Forces

Van der Waal's Forces, weak – caused by random movement of electrons which result in temporary dipoles – occurs between non-polar molecules and give low melting and boiling points. Are also found between plates of graphite. **O₂ boiling point is higher than H₂ because the molecule is bigger so there is more chance of temporary dipoles forming – not because O₂ is heavier**

Dipole-Dipole Forces $\delta +$ of a polar molecule attracts $\delta -$ of another. Results in higher melting and boiling points than similar mass non-polar compounds e.g. ethane C₂H₆ [low] and HCHO methanal [high due to polar carbonyl group C=O]. Stronger than van der Waal's because they are permanent dipoles

Hydrogen Bonding is a special type of dipole-dipole and occurs **when H is bonded to a more electronegative element such as F, N, or O**. It is found between water molecules or where there is a large degree of polarity in small molecules. It causes much higher boiling points and surface tension. H₂S has a lower melting and boiling point than water even though its molecule is bigger, as the effect of polarity is not as great as it is in water.

k 2.8 Oxidation Numbers

Oxidation is Loss of Electrons - **Reduction** is Gain of Electrons – [OIL RIG]

Oxidizing Agents – cause oxidation e.g. O₂, Cl₂, Cr₂O₇²⁻, MnO₄¹⁻

Reducing Agents – cause reduction e.g. hydrogen, carbon and carbon monoxide

Oxidation Numbers

The number of electrons an atom appears to have gained or lost when distributed according to certain arbitrary rules.

Rules

1. Elements = 0
2. Simple ion = charge on ion
3. Complex ions \Rightarrow sum of oxidation numbers = charge on ion
4. Compounds \Rightarrow sum of oxidation numbers = 0
5. Alkali Metals in compounds = +1
6. Alkaline Earth Metals in compounds = +2
7. Halogens in compounds = -1
8. H = +1 except in metal hydrides [NaH and KH] where it is -1
9. Oxygen = -2 except in peroxides [H₂O₂] where it is -1 and when bonded to a more electronegative element e.g. OF₂ where it is +2

Oxidation is an **increase in Oxidation Number** e.g. -3 to -2 or +2 to +3

Reduction is a **decrease in Oxidation Number** e.g. from -2 to -3 or +1 to 0

Calculation of Oxidation Numbers (exclude peroxides, except for H₂O₂).

Apply rules above and the element that is not covered is then worked out

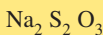
Assign oxidation numbers to all elements in Na₂S₂O₃

Na = +1 [Rule 5]

O = -2 [Rule 9]

S not covered so work it out

Write Formula



Assign individual oxidation number +1 ? -2

Assign totals of oxidation numbers +2 ? -6

Remember total = 0 [Rule 4]

S₂ must = +4

Each S = +4/2 = +2

Always put in pluses and minuses