2.1.1 Atomic structure and isotopes

Definitions

Term	Definition
Isotopes	Atoms of the same element with the same number of **protons and electrons** and **different numbers of neutrons and different masses**.
Proton number / atomic number	The number of **protons** in the nucleus of an atom.
Nucleon number / mass number	The number of **protons and neutrons** in the nucleus of an atom.
Relative isotopic mass / A	The mass of an isotope of an element compared to 1/12th of the mass of an carbon-12 atom.
Relative atomic mass / A_r / A_R	The **weighted mean** mass of an atom of an element compared to 1/12 of the mass of an atom of carbon-12.
Cation	A **positively charged** ion with fewer electrons than protons.
Anion	A **negatively charged** ion with more electrons than protons.

Concepts

- Properties of isotopes
    - Same chemical reactions
        - Same electron configuration & the same number of protons
        - Number of neutrons has no effect on reactions of an element
    - Small differences in physical properties
        - Higher mass isotopes = higher melting and boiling point + higher density
- Mass and charge of sub-atomic particles

Particle	Relative charge	Relative mass
Proton / p^+	1+	1
Neutron / n	0	1
Electron / e^-	1-	1/1836

- Determining relative atomic mass and relative isotopic mass (for ions with single charges)
    - Mass spectrometer
    - Records abundance of ions of different isotopes and their mass-to-charge ratio (m/z ratio)
    - Value of relative isotopic mass can be worked out from m/z ratio and hence relative atomic mass

2.1.2 Compounds, formulae and equations

Definitions

Term	Definition
Binary compounds	Compounds that contains **two elements** only.
Diatomic molecules	Molecules composed of **two atoms** only.
Polyatomic ions	Ion containing more than one atoms.

Concepts

- Anions to know

Ion	Formula
Nitrate	NO_3^-
Carbonate	CO_3^{2-}
Sulfate	SO_4^{2-}
Hydroxide	OH^-

- Cations to know

Ion	Formula
Ammonium	NH_4^+
Zinc ion	Zn^{2+}
Silver ion	Ag^+

- Writing ionic equations
  - We can only dissociate the **aqueous compounds**
  - Split all chemicals into ions
  - Cancel out spectator ions
  - ★ No aqueous compound = no ionic equation
- Solubility

Solubility	Compounds
Soluble in water	• All common sodium, potassium and ammonium salts (also their carbonate and hydroxide salt) • All nitrates • Most common chlorides • Most common sulfates
Insoluble in water	• **Silver chloride, lead chloride** • **Lead sulfate, barium sulfate, calcium sulfate, strontium sulfate** • Most common carbonates • Most common hydroxides

2.1.3 Amount of substance

Definitions

Term	Definition
Mole	A mole is the amount of a substance that contains the Avogadro number of elementary particles.
Molar mass / M	The mass in grams in each mole of the substance, measured in g mol^{-1}.
Hydrated	A crystalline compound that contains water (e.g. $CuSO_4 \cdot 5H_2O_{(s)}$).
Anhydrous	A crystalline compound containing no water (e.g. $CuSO_{4(s)}$).
Water of crystallisation	Water molecules that form part of the crystalline structure of a compound (e.g. H_2O in $CuSO_4 \cdot 5H_2O_{(s)}$).
Stoichiometry	The relative quantities of substances in a reaction.
Standard solution	A solution of known concentration.
Limiting reagent	The reactant that is not in excess and will be used up in the reaction.

Amount of substance

- Amount of substance
  - Symbol n
  - Measured in moles (symbol mol)
  - * Always use **decimals (not fractions)** in **every step** of a calculation
- Avogadro constant / N_A
  - $6.02 \times 10^{23}\ \text{mol}^{-1}$
  - The number of particles per mole
- Concentration (c)
  - Unit = mol dm^{-3} (aka molar / M) or g dm^{-3}
  - mol dm^{-3}: $c = \frac{n}{V} = \frac{\text{number of moles}}{\text{volume (in dm}^{-3})}$
  - g dm^{-3}: $c = \frac{\text{mass (in g)}}{\text{volume (in dm}^{-3})}$
  - $\text{Concentration in mol dm}^{-3} = \frac{\text{concentration in g dm}^{-3}}{M_r}$
- Room temperature and pressure (RTP)
  - Temp = **20 °C / 293 K**
  - pressure = 1 atm or 1.01×10^5 Pa
- Standard temperature and pressure (STP)
  - Temp = 0 °C / 273 K
  - pressure = 1 atm or 1.01×10^5 Pa
- Molar gas volume / V_m
  - The volume per mole of gas at a stated temperature and pressure
  - Under RTP: 1 mol = **24 dm^3 = 24,000 cm^3**
  - Under STP: 1 mol = 22.4 dm^3 = 22,400 cm^3
- Experimental techniques to measure the amount of substances
  -

Variable measured	Method
Mass	• Use a digital mass balance • Choose a balance with a suitable resolution for the experiment
Volume of solution	• Use a measuring cylinder • Standard solution: use volumetric flask
Gas produced	• Use a gas syringe / measure mass lost on a balance and calculate the

	number of moles of gas produced

- Percentages yield
    - $\text{Percentage yield} = \dfrac{\text{actual yield}}{\text{thereotical yield}} \times 100\%$
    - Actual yield: the amount of the product obtained from a reaction
    - Theoretical yield: the yield resulting from complete conversion of reactants into products
    - Reasons for < 100% percentage yield
        - Reaction did not go to completion (reversible reaction)
        - Side reactions may have taken place along the main reaction
        - Purification of the product may result in the loss of some products
- Atom economy
    - $\text{Atom economy} = \dfrac{\text{sum of masses of useful product(s)}}{\text{sum of masses of all products or reactants}} \times 100\%$

      $= \dfrac{\text{sum of molar masses of useful products}}{\text{sum of molar masses of all products or reactants}} \times 100\%$

2.1.4 Acids

Definitions

Term	**Definition**
Acid	When dissolved in water an acid releases H^+ ions (proton) into the solution.
Bases	React with acid by **accepting H^+ ions (protons)** and neutralising the acid to form a salt.
Alkalis	Bases that dissolve in water and release OH^- ions into the solution.
Salt	Formed when the H^+ in an acid is replaced by a positive ion

Common chemicals

- Common acids
  - HCl
  - H_2SO_4
  - HNO_3
  - CH_3COOH
- Common bases
  - Carbonates
  - Hydrogencarbonates
  - Metal oxides
  - Metal hydroxides
  - **Ammonia (accept H^+ and form NH4$^+$ ions)**
- Common alkalis
  - NaOH
  - KOH
  - **NH_3**

Acid-base reactions

- Strong and weak acid
  - Both release H^+ ions / H^+ donor in aqueous solutions
  - Strong acid
    - **Completely dissociates** (> 99%) in aqueous solutions / releases all hydrogen atoms as H^+ ions
    - e.g. $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$
  - Weak acid
    - **Partially dissociates** (< 1%) in aqueous solutions / only releases a portion of available hydrogen atoms as H^+ ions
    - e.g. $CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$
- Neutralisation
  - The reaction of acids with bases (including carbonates, metal oxides and alkalis) to form salts
  - **Ionic equation: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$**

Experimental techniques

- Preparing standard solution
  - Solid weighed accurately **using a digital balance with 2 dp or more**
  - Dissolve solid in a beaker using distilled water
    - Use less distilled water than needed to fill the volumetric flask to the mark
  - Transfer the solution to (250 cm^3) volumetric flask
    - Rinse the beaker and transfer washings to the flask so the last traces of the solution is transferred to the volumetric flask
  - Volumetric flask is filled to the graduation line

        - Add distilled water a drop at a time using a dropping pipette
        - Keep adding until the **bottom of the meniscus** lines up exactly with the mark
    - Mix the solution thoroughly
        - Volumetric flask is sealed with a stopper and inverted several times
- Titration
    - Add measured volume of one solution to conical flask using pipette
        - Typical tolerances: 10 cm^3: ± 0.04 cm^3, 25 cm^3: ± 0.04 cm^3, 50 cm^3: ± 0.10 cm^3
    - Add other solution to burette, record initial reading
    - Add a few drops of indicator to conical flask
        -

Formula	**Acidic**	**Neutral**	**Alkaline**
Phenolphthalein	Colourless	Colourless	Pink
Methyl orange	Red	Yellow	Yellow

-
    - Run solution from burette into conical flask until it reaches the end point (indicated by colour change)
        - Swirl the flask while the solution is added
    - Record final reading
    - Titre = final reading - initial reading
    - First titre carried out quickly to get approximate titre
    - Repeat accurately by adding solution dropwise as the end point is approached
    - Carry out until two accurate titres are concordant (within 0.1 cm^3)
    - Only use concordant results for calculating the mean titre

2.1.5 Redox

Definitions

Term	Definition
Redox reactions	A reaction involving reduction and oxidation.
Oxidising agent	A reagent that **accepts / takes in** electrons / oxidises another species / is reduced
Reducing agent	A reagent that **donates / gives out** electrons / reduces another species / is oxidised
Oxidation number	A measure of the number of electrons that an atom uses to bond with atoms of another element. Oxidation numbers are derived from a set of rules.

Oxidation number

- Oxidation number (oxidation state) rules
    - Elements
        - **Always 0**
        - Any bonding is to atoms of the same element in pure elements
    - Compound and ions
        - Each atom in a compound has an oxidation number
        - Sign is placed before the number
        - Sum of oxidation numbers in a compound / ion = total charge
- Fixed oxidation numbers

Combined element	Oxidation number
O (normally)	-2
H (normally)	+1
F	-1
Group 1	+1
Group 2	+2
Group 3	+3

- Oxidation number for special cases

Combined element	Oxidation number
H in metal hydrides (e.g. NaH, CaH_2)	-1
O in peroxide ions (O_2^{2-})	-1
O bonded to F (e.g. F_2O)	+2

- Roman numerals in chemical names
    - Show oxidation number without sign
    - Nitrate = assume to be NO_3^-
    - Sulfate = assume to be SO_4^{2-}
    - e.g. chlorate(I) = ClO^-

Redox reactions

- How to identify redox reactions
    - Identifying oxidation
        - Gain of oxygen
        - Loss of hydrogen
        - Loss of electrons
        - Increase in oxidation number

    - Identifying reduction
        - Loss of oxygen
        - Gain of hydrogen
        - Gain of electrons
        - Decrease in oxidation number
    - * Oxidation and reduction always **happen together**
- Redox reaction of acids
    - Metal + acid → salt + hydrogen
    - Metal oxidised (oxidation number increases from 0 to …)
    - Hydrogen in acid reduced (oxidation number decreases from +1 to 0)
    - (Iron is normally Fe^{2+} in redox reactions)

2.2.1 Electron structure

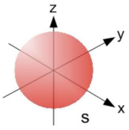
Definitions

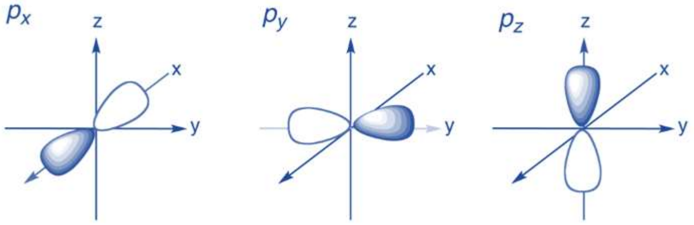
Term	**Definition**
Atomic orbitals	A region around the nucleus that can hold up to two electrons with opposite spins.
Sub-shells	A group of orbitals of the same type within a shell.
Electron configuration	A shorthand method for showing how electrons occupy sub-shells in an atom.

Atomic orbitals

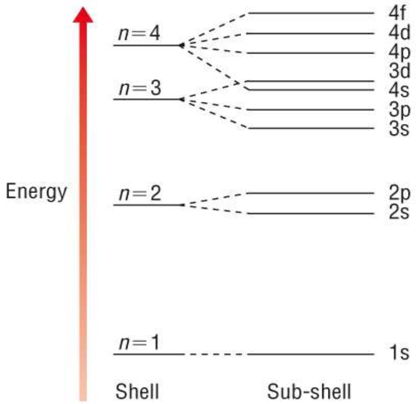
- Number of electrons that can fit in each shell

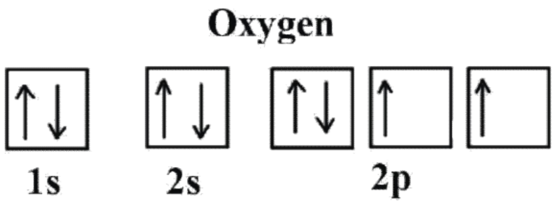
Shell number n	**Number of electrons ($= 2n^2$)**
1	2
2	8
3	18
4	32

- Types of orbitals
  - s-orbitals
    - **Spherical** shape
    - Each shell from $n = 1$ contains 1 s-orbital
    - Greater shell number n = greater radius of its s-orbital
    - 

  - p-orbitals
    - **Dumb-bell** shape
    - Three separate p-orbitals at right angles to one another: p_x, p_y, p_z
    - Each shell from $n = 2$ contains 3 p-orbitals
    - Greater shell number n = its p-orbital is further from the nucleus
    - 

  - d-orbitals
    - Each shell from $n = 3$ contains 5 d-orbitals
  - f-orbitals
    - Each shell from $n = 4$ contains 7 f-orbitals
- Filling of sub shells and orbitals
  - Sub shells fill in the order of **increasing energy** (1s → 2s → 2p …)

    - 

    - **4s is filled before 3d**
- Inside each sub-shell
    - The orbitals all have the same energy within a sub-shell
    - One electron occupies each orbital before pairing starts to prevent any repulsion between paired electrons
    - Opposite spins within each orbital (one ↑ and one ↓) as both electrons are negatively charged and repel one another so this minimises repulsion
    - The paired electrons are easier to remove than the unpaired ones
    - 

- More than one orbital within a sub-shell = the orbitals are degenerate (all the same)

Electron configurations

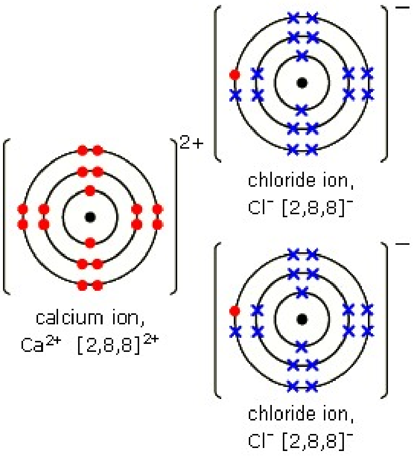
- Writing electron configuration of atoms
    - e.g. Li = $1s^22s^1$, F = $1s^22s^12p^5$
    - Shorthand notation: in terms of the previous noble gas + outer electron sub-shells
        - e.g. Krypton = $1s^22s^22p^63s^23p^63d^{10}4s^24p^6$ = [Ar] $3d^{10}4s^24p^6$
        - **Always show outer shell electrons**
    - (4s can be written before 3d)
- Exceptions of atom electron configurations
    - Chromium: [Ar]$\mathbf{3d^54s^1}$
    - Copper: [Ar]$\mathbf{3d^{10}4s^1}$
    - Chromium and copper do not follow the expected pattern
        - Half-filled / fully filled d sub-shell gives **additional stability**
- Electron configuration of ions
    - The highest energy subshell gain or lose electrons
    - d-block elements
        - 4s is at a lower energy level than 3d sub-shell so it is filled first
        - Once filled the energy level of 3d falls below 4s so 4s empties before 3d

2.2.2 Bonding and structure

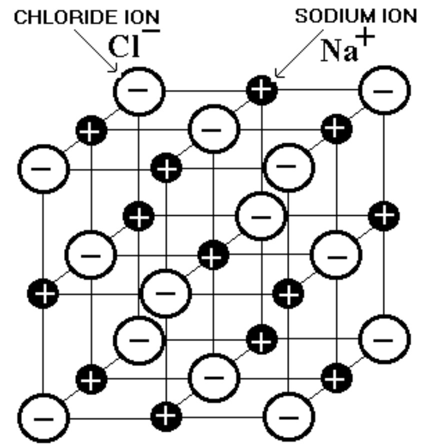
Definitions

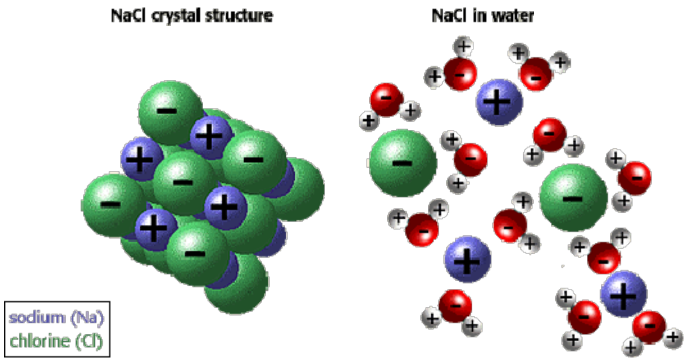
Term	Definition
Polarity	There is an uneven distribution of electrons.
Polar molecule	A molecule with an overall dipole, having taken into account any dipole across bonds and the shape of the molecule.
Bond angle	The angle between 2 bonds in a molecule.
Electronegativity	A measure of the attraction of a bonded atom for the shared pair of electrons in a covalent bond.
Dipole	A separation in electrical charge so that one atom of a polar covalent bond, or one end of a polar molecule, has a small positive charge, δ+, and the other has a small negative charge, δ-.
Intermolecular forces	An attractive force between molecules. Can be London forces, permanent dipole-dipole interactions or hydrogen bonding.
Giant ionic lattice	A three dimensional structure of oppositely charged ions, bonded together by strong ionic bonds.
Simple molecular lattice	A three dimensional structure of molecules, bonded together by weak intermolecular forces.

Ionic compounds

- Ionic bonding
  - The strong electrostatic attraction between positive and negative ions
  - Occurs between atoms that have a **big difference in electronegativity** (usually metals and non-metals)
- Dot and cross diagram
  - 

- Structure of ionic compounds
  - Each ion attracts oppositely charged ions strongly in **all directions**
  - Forms a **giant lattice structure** containing many ions
  - Every positive ion is surrounded by negative ions; every negative ion is surrounded by positive ions
  - Regular arrangement of ions = regular, predictable shapes depending on size of ions
    - NaCl = always cubic
    - $MgSO_4$, $CuSO_2$, etc. = not cubic

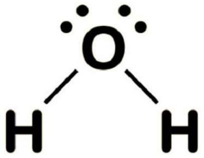
    - 

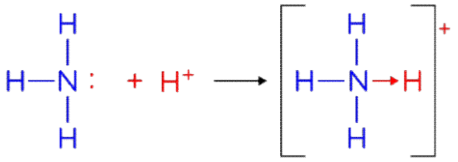
SODIUM CHLORIDE LATTICE STRUCTURE
- Properties
    - High melting and boiling points
        - **Strong** electrostatic forces of attraction between oppositely charged ions in the **giant ionic lattice**
        - A high temperature is needed to provide the large amount of energy needed to overcome the strong electrostatic attraction
        - Higher for ions with greater ionic charges due to stronger attraction between ions
        - * Ionic substances with complex ions: can decompose before they reach melting point
    - Dissolve in polar solvents (e.g. water)
        - Salvation / hydration
            - Polar solvent molecules attract the ions in the lattice
            - Anions to positive dipole, cations to negative dipole
            - Energy is released in the process
            - Molecules break down the lattice + surround each ion in solution
        - Strong ionic bond must be overcome for the substance to dissolve
            - Energy to do this is supplied by solvating / hydrating the ions / absorbed from the surroundings
        - If the compound is made of ions with strong charges the attraction can be too strong to be broken down → less soluble
        - Solubility depends on
            - Relative strengths of attractions within the lattice
            - Attraction between ions + water molecules
        - 

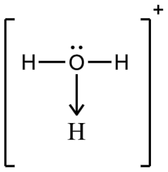
    - Does not conduct electricity in solid state
        - Ions held in fixed positions
        - No mobile charge carriers
    - Conduct electricity when dissolved or molten
        - The solid ionic lattice breaks down
        - Ions are free to move as mobile charge carriers

Covalent compounds

- Covalent bonding
    - The **strong** electrostatic attraction between a shared pair of electrons and the nuclei of the

bonded atoms
    - Occur between atoms that have a **small difference in electronegativity**
    - Formed when the atomics orbitals of the two atoms overlap and combine to form a new orbital (still holding 2 electrons max)
- Covalent substance structure
    - Attraction is **localised** between the shared pair of electrons and the nuclei of the bonded atoms only
    - Results in **small molecules** consisting two or more atoms (simple molecular lattice if solid)
- Lone pairs / non-bonding pair
    - **Paired** electrons that are **not shared**
    - 

- Multiple covalent bond
    - The electrostatic attraction is between **two / three shared pairs** of electrons and the nuclei of the bonded atoms
- Dative covalent bonds / coordinate bonds
    - The shared pair of electrons is supplied by one of the bonding atoms only
    - The shared pair is originally a lone pair of electrons on one of the bonded atoms
    - e.g. ammonia molecule donates its lone pair of electrons to a H^+ ion to form NH_4^+ ion
    - 

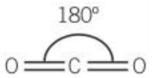
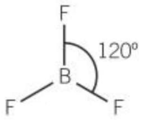


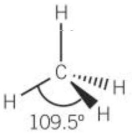
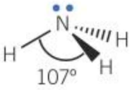
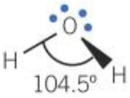
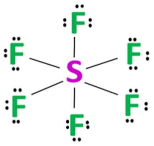
Hydronium ion

- Exceptions
    - Boron
        - Only 3 outer shell electrons can be paired → only forms 3 pairs
        - e.g. BF_3 only has 6 electrons around the boron atom
    - Phosphorus, sulfur, chlorine (expansion of the octet)
        - Outer electrons are in $n = 3$ shell which can hold up to 18 electrons
        - More electrons are available for bonding (can have more than 8 electrons in outer shell e.g. ClF_7)
- Average bond enthalpy
    - A measurement of covalent bond strength
    - Higher average bond enthalpy = stronger covalent bond

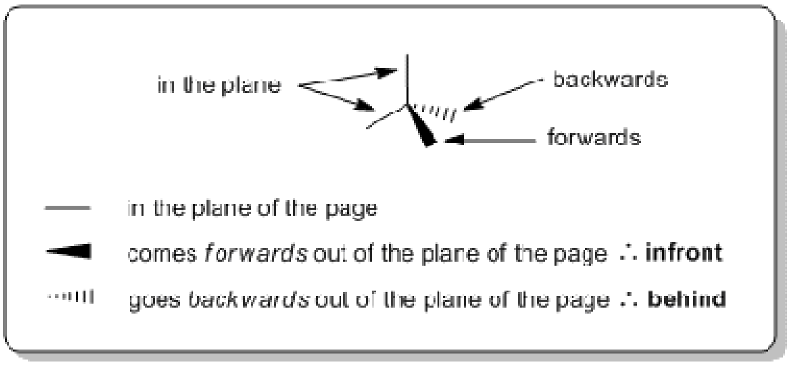
Molecule structures

- Molecule shape and bond angles
    - Electron pairs around the central atom repel each other as far apart as possible
    - Lone pairs repel more strongly than bonded pairs
    - They repel bonded pairs slightly closer together → reducing the bond angle
        - 4 pairs / regions: bond angle reduced by approx. 2.5° per lone pair
    - The greater the number of electron pairs the smaller the bond angle
    - (Mention the number of bonded pairs and lone pairs)

Electron pairs	Bonded pairs	Lone pairs	Name	Shape and bond angle
2	**2**	**0**	Linear	
3	**3**	**0**	Trigonal planar	

4	**4**	**0**	Tetrahedral	
4	**3**	**1**	Pyramidal	
4	**2**	**2**	Non-linear	
6	**6**	**0**	Octahedral	 **Remember SF_6 as an example of octahedral arrangement**

- Bond representations



in the plane of the page

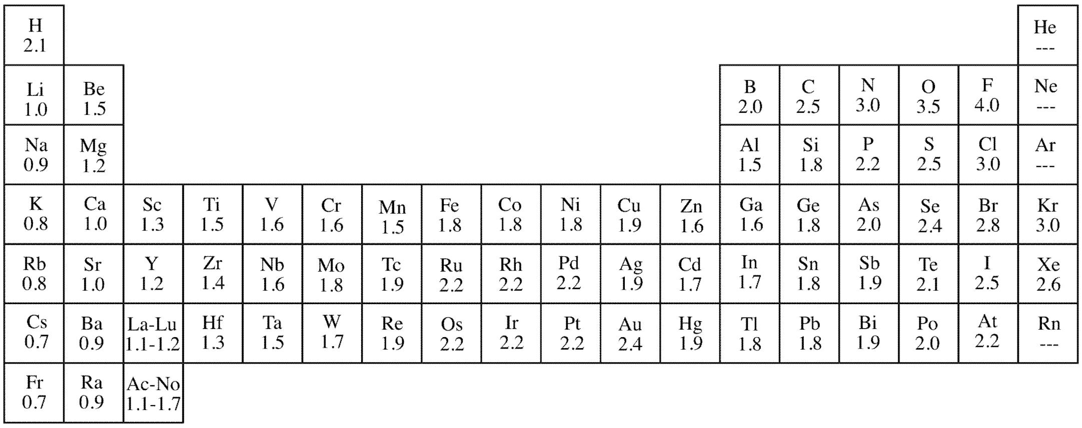
comes *forwards* out of the plane of the page ∴ **infront**

goes *backwards* out of the plane of the page ∴ **behind**

- Predicting molecular shape of ions
  - Ions with oxygen
    - Add an electron to an oxygen atom for every negative charge
    - Bonds can be single or double bond
  - Cations with hydrogen
    - Add positive charge by H^+ ions
  - Others
    - Add / remove electrons from the central atom to match the charge
    - Pair up surrounding atoms and count bonded pairs and lone pairs

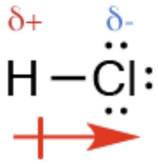
Electronegativity and bond polarity

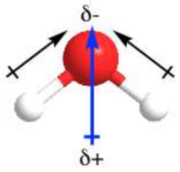
- Electronegativity
  - A measure of the attraction of a bonded atom for the pair of electrons in a covalent bond
  - Measured using Pauling electronegativity values (higher = attract bonded electrons more strongly)



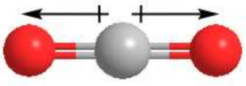
H 2.1																	He ---
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne ---
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.2	S 2.5	Cl 3.0	Ar ---
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr 3.0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe 2.6
Cs 0.7	Ba 0.9	La-Lu 1.1-1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn ---
Fr 0.7	Ra 0.9	Ac-No 1.1-1.7															

- Important electronegativity values

Element	Electronegativity
Fluorine	4.0
Oxygen	3.5
Chlorine	3.0
Nitrogen	3.0
Carbon	2.5
Hydrogen	2.1
- Electronegativity pattern
  - Electronegativity increases across the periodic table and up the periodic table
  - Across the periodic table the nuclear charge increases and the atomic radius decreases → attraction increases
  - Atomic radius increases down the table + increased shielding → attraction decreases
  - **Fluorine is the most electronegative element**
- Predicting type of bond
Bond type	Electronegativity difference
Covalent	0
Polar covalent	0 - 1.8
Ionic	> 1.8
- Pure covalent bond
  - **Non-polar** bond
  - The bonded electron pair shared equally between the bonded atoms
  - Happen when:
    - The bonded atoms are the same
    - The bonded atoms have same / similar electronegativity (e.g. carbon and hydrogen)
- Polar covalent bonds
  - Polar bond
  - The bonded electron pair is shared unequally between the bonded atom
  - When the bonded atoms are different and have different electronegativity values
  - Form permanent dipole: small positive (δ+) / negative charge (δ-) on the two bonded atoms, oppositely charged on each
  - 

- Polar / non-polar molecules
  - Polar molecules require polar bonds with dipoles that **do not cancel out** due to their direction
  - If the molecule is symmetrical in 3D **and** the polarities cancel out the molecule is non-polar
  - e.g. overall dipole in H_2O, cancels out in CO_2 and methane
  - 

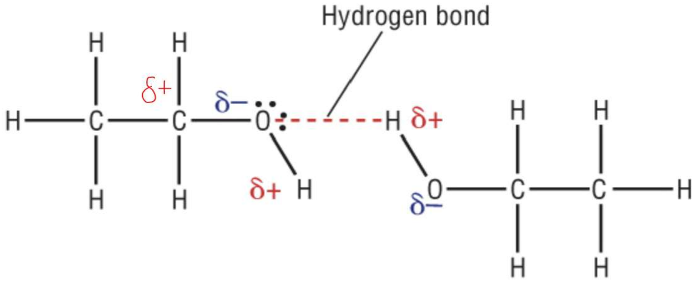
    H_2O, bend shape
    polar

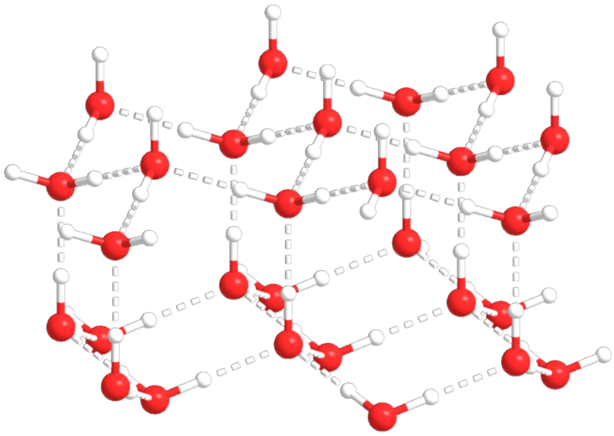
    CO_2, linear shape
    non-polar

Intermolecular forces

- Types of intermolecular forces
  - London forces / induced dipole-dipole **interactions** / dispersion forces (van der Waals' forces)
  - Permanent dipole-dipole **interactions** (van der Waals' forces)
  - Hydrogen bonding (a special type of permanent dipole-dipole interactions)

- London forces / induced dipole-dipole interactions
    - Exist between all molecules, only temporary
    - Creation
        - Movement of electrons produces changing dipole in a molecule
        - There is an uneven distribution of electrons which causes an instantaneous dipole
        - At any instant an instantaneous dipole will exist but position shifts constantly
        - Instantaneous dipole induces a dipole on a neighbouring molecule
        - The induced dipole induces dipoles on further molecules, they then attract one another
        - Electrons move and the dipole is gone and the attraction disappears and re-appears elsewhere
    - Very weak - the weakest of all intermolecular forces
        - The **only intermolecular** force that exist between non-polar simple covalent molecules
        - They have very low melting and boiling points (normally gases / volatile liquids under room temperature)
- London forces strength pattern
    - Stronger down the periodic table
    - More electrons in each molecule
    - Larger instantaneous + induced dipoles
    - Greater induced dipole-dipole interactions
    - Stronger electrostatic attractive forces between molecules
- Permanent dipole-dipole interactions
    - Act between **permanent dipoles** in polar molecules
    - Permanent dipole-dipole interactions is much stronger than London forces
        - Polar molecules have both London forces and permanent dipole-dipole interactions
        - Mean melting and boiling point of polar molecules is much greater
- Hydrogen bond
    - Occurs when a highly de-shielded hydrogen in one molecule is attracted to a lone pair on a **N, O or F** atom in another molecule
        - Highly de-shielded hydrogen: a hydrogen atom directly bonded to an **very electronegative** atom (**N, O or F**)
    - A lone pair of electrons on a N/O/F atom forms bond with the $\delta+$ of a hydrogen atom in a different molecule
    - Strongest type of intermolecular attractions
        - Molecules with hydrogen bonds between them have relatively high melting and boiling points
        - More hydrogen bond = greater increase in melting and boiling points
- Hydrogen bond diagram
    - Shown by a dotted line between lone pair + hydrogen atom
    - Should have a **straight line through all the atoms involved**
    - 

- Anomalous properties of H_2O
    - Solid less dense than liquid
        - Each water molecule can form **4 hydrogen bonds** (2 lone pairs + 2 hydrogen atoms in O-H bond)
        - Hydrogen bonds extend outwards and **hold water molecules apart** in an **open lattice structure**
        - Water molecules form an **tetrahedral lattice** full of holes
        - Water molecules in ice are further apart than in water
        - (Bond angle for hydrogen involved in hydrogen bond is 180°)

        - 
    - Relatively high melting and boiling points
        - Has hydrogen bond (much stronger) + PDD + IDD between molecules
        - A large quantity of energy is needed to break hydrogen bonds
        - When ice lattice breaks the arrangement of hydrogen bonds is broken
        - When water boils hydrogen bonds break completely

Simple molecular substances

- Simple molecular substance
    - Made of simple molecules (small units with a definite number of atoms)
    - Simple molecules form a regular structure called a simple molecular lattice in the solid state
        - Shape of the lattice is determined by the packing pattern of the molecules
        - Molecules held together by weak intermolecular forces
        - Atoms within each molecule bonded together by strong covalent bonds (**intra**molecular forces)
- Properties of non-polar simple molecular substances
    - Low melting and boiling point
        - Only weak London forces between molecules
        - Weak intermolecular forces can be broken by energy present at low temperatures
    - Solubility
        - Non-polar solvent
            - Tend to be soluble
            - Intermolecular forces form between the molecules and the solvent
            - Weaken the intermolecular forces in the simple molecular lattice
            - The intermolecular forces break and the compound dissolves
        - Polar solvent
            - Tend to be insoluble
            - Little interaction between the lattice and the solvent molecules
            - **Intermolecular bonding in solvent** is too strong to be broken
    - Non-conductors of electricity
        - No free electrons or mobile ions in simple molecular substances, cannot complete a circuit
        - No mobile charge carriers
- Properties of polar simple covalent substance
    - Higher melting and boiling point
        - Permanent dipole-dipole interactions are stronger than London forces
        - More energy is required to break them
    - Solubility
        - Hard to predict
        - Usually soluble in polar solvents: polar solute/solvent molecules can attract each other
        - If they contain a significant non-polar region they may also be soluble in non-polar solvents
        - Solubility depends on strength of dipole
    - Electrical conductivity
        - Can conduct electricity if they can ionise and form mobile ions

    - Otherwise same as above