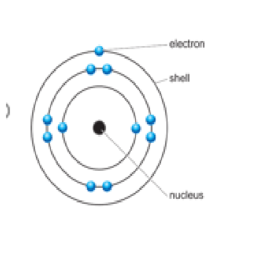
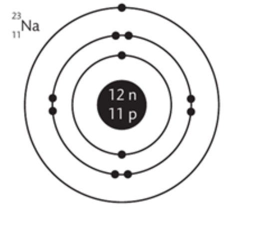
Applied Science Unit 1 Chemistry Knowledge Organiser

Electronic orbitals

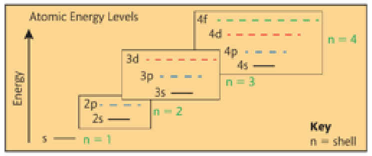


Bohr theory



Aufbau principle

Electronic configuration: The distribution of electrons in an atom/molecule.

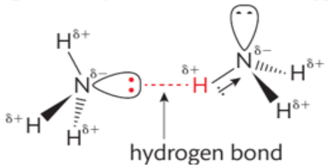


Intermolecular forces

Dipole: Separation of charges within a covalent molecule.
Van der Waals forces: all intermolecular attractions are van der Waals forces.
Dipole-dipole forces: Permanent forces between polar molecules

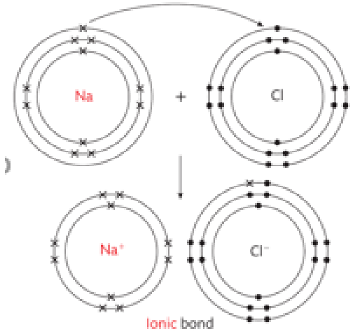
Figure 1.16: Dipole-dipole forces

Hydrogen bonding: Strongest form of intermolecular force. Hydrogen bonds will form when compounds have hydrogen directly bonded to **fluorine**, **oxygen** or **nitrogen**



Ionic bonding

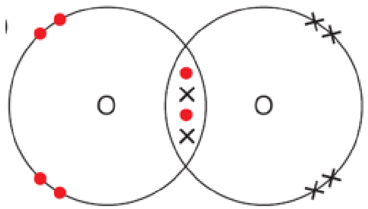
Occurs when an atom of an element **loses** one or more electrons and **donates** it to an atom of a different element.



Electrostatic attraction: The force experienced by oppositely charged particles. Holds particles strongly together.
Giant ionic lattice: a regular arrangement of **positive ions** and **negative cations** e.g. NaCl.

Covalent bonding

Strong **electrostatic attraction** between two **non-metal** nuclei and the shared pair(s) of electrons between them.



Lone pair: a non-binding pair of electrons.

Organic compound – a compound that contains one or more carbons in a carbon chain.

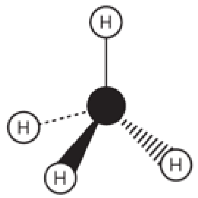
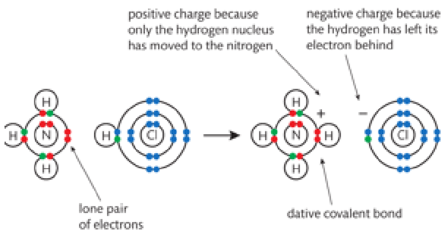


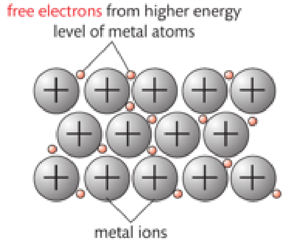
Figure 1.9: Tetrahedral structure of methane



Each carbon atom has **covalently** bonded to 4 hydrogen atoms. **Tetrahedral structure** due to **negative electron pair repulsion.**

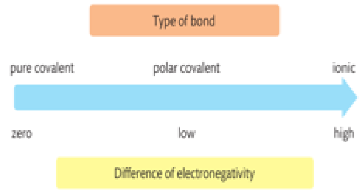
Metallic bonding

Metallic bonding is caused because the electrons in the highest energy level of a metal atom has the ability to become **delocalised**.



Delocalised electrons: electrons that are free to move. They are not associated with a single atom or covalent bond.

Electronegativity: The tendency of an atom to attract a binding pair of electrons.
Non-polar: A molecule where the electrons are distributed **evenly** throughout the molecule.
Polar molecule: A molecule with a partial positive charge in one part of the molecule and similar negative charge in another part due to uneven electron distribution



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Empirical formula

Shows the ratio between elements in a chemical compound.

Step 1: Divide each mass of each element present in a compound by its molar mass to get its molar ratio.
Step 2: Divide the answer for each element by the smallest molar ratio calculated.
Step 3: If the answers are not whole numbers, multiply them by the same number to get whole numbers.

Molecular formula

Used for simple molecules. To work out the molecular formula you need to know the **empirical formula** and the **relative molecular mass.** x

E.g. Compound has the empirical formula CH_2
Empirical formula mass = 12 + (1 x 2)
Relative molecular mass = 42
Molecular formula = 42/12 = 3
Molecular formula = CH_2 x 3 = C_3H_6

Percentage yield

Percentage yield is the actual mass compared to the theoretical mass.
Theoretical mass: The **expected** amount of product from a reaction calculated from the balanced equation.

The formula for calculating percentage yield is:

$$\text{Percentage yield} = \frac{\text{actual number of moles}}{\text{expected number of moles}} \times 100\%$$

It can also be calculated as:

$$\text{Percentage yield} = \frac{\text{actual mass}}{\text{theoretical mass}} \times 100\%$$

Reacting quantities

Titration: A method of volumetric analysis used to calculate concentration of a solution.
Solution: a liquid mixture where a solute is dissolved in a solvent.
Solute: The substance dissolved in a solvent to form a solution.
Solvent: a liquid that dissolves another substance.

The number of moles of **solute** in a given volume of **solvent** tells you how **concentrated** the solution is.

1 mole of solute dissolved in 1 cubic decimetre of solution is written as:

1 mol dm^{-3} or 1M for short

1 mole of HCl has a mass of = 36.5g
36.5g of HCl in 1 dm^{-3} has a concentration of 1 mol dm^{-3} or 1M

Example:

How many moles of hydrochloric acid are there in 100 cm^3 of 1M hydrochloric acid solution?

Number of moles (N) = molarity (C) × volume of solution (V) (dm^3)
$N = CV$

The volume is given in cm^3 so this needs to be converted into dm^3 by dividing by 1000. (Remember 1 dm^3 = 1000 cm^3)

$$\text{Number of moles} = \frac{100}{1000} \times 1$$
$$= 0.1 \text{ mol}$$

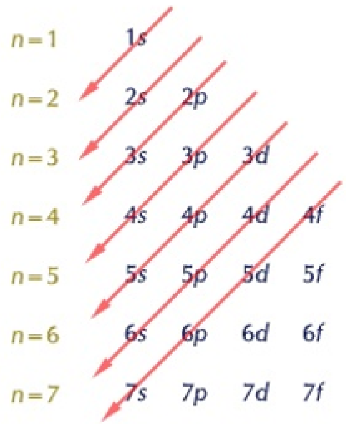
Periods 1,2,3 and 4

Period	Characteristics
1	Contains hydrogen and helium. Both are gases. The electrons in these two elements fill the 1s orbital. Helium only has two electrons and, chemically, is unreactive. Hydrogen readily loses or gains an electron, and so can behave chemically as both a group 1 and a group 7 element. Hydrogen can form compounds with most elements and is the most abundant chemical element in the universe.
2	Contains eight elements: lithium, beryllium, boron, carbon, nitrogen, oxygen, fluorine and neon. The outer electrons in these elements fill the 2s and 2p orbitals. Nitrogen, oxygen and fluorine can all form diatomic molecules. Neon is a noble gas. Carbon is a giant molecular structure.
3	Contains eight elements: sodium, magnesium, aluminium, silicon, phosphorus, sulfur, chlorine and agon. The outer electrons in these elements fill the 3s and 3p orbitals.
4	Contains 18 elements, from potassium to krypton. The first row of the transition elements is in this period. The outer electrons on these elements fill the 4s, 4p and 3d orbitals.

Groups: s block, p block, d block

Element blocks in the periodic table are named for the orbitals that the **highest energy electrons** are in for that set of elements. Group 1 and 2 of the periodic table are in **s block**. Group 3 to 7 and group 0 make up **p block**. Transition metals are in the **d block.**

For example:
Carbon has electronic structure of $1s^2\ 2s^2\ 2p^2$
The highest energy electron in carbon is in a p orbital and therefore carbon is a p block element



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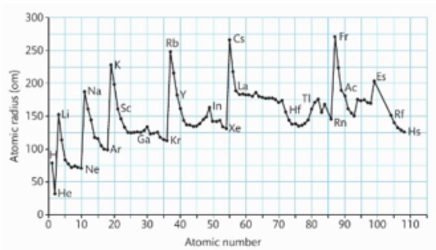
Atomic radius

Atomic radius **decreases** across the **period** from left to right. Across the group more protons and electrons are added, but they are added to the same s and p sub-shells so the size does not increase. As you go down a **group**, the atomic radii **increases**.



Ionic radius

Cations: ions with a positive charge.

Anions: ions with negative charge.

Isoelectronic: having the same number of electrons.

Extra electrons are added to extra shells as you go down the group, therefore giving a larger size.

As you go across a period, the **cations** all have the same electronic structure. They are **isoelectronic**. The nuclear charge increases – this pulls the electrons more strongly thus ionic radii of cations **decrease** as you go across the period.

Anions have a larger radius than the corresponding atom because there is more **repulsion** between the extra electrons. As you go across the period, they are **isoelectronic**. The number of protons also increases, so the ionic radius also **decreases** as you go across the period.

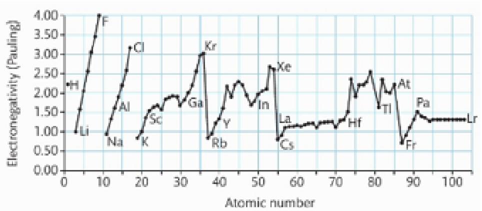
Electronegativity

Electronegativity is a measure of the tendency of an atom to **attract** a bonding pair of electrons. It **increases** as you go across a period. It **decreases** as you go down a group.

Electronegativity depends on the number of **protons**, the **distance** from the nucleus of the bonding pair and how much **shielding** there is from inner electrons.

As you go across the **period**, bonding pair of electrons will be shielded by the **same** number of electrons. However, the **number** of protons will **increase**. Group 7 are more electronegative than group 1.

As you go down the **group**, there is more **shielding**. Bonding pair of electrons are further from the nucleus. Electronegativity **decreases**.

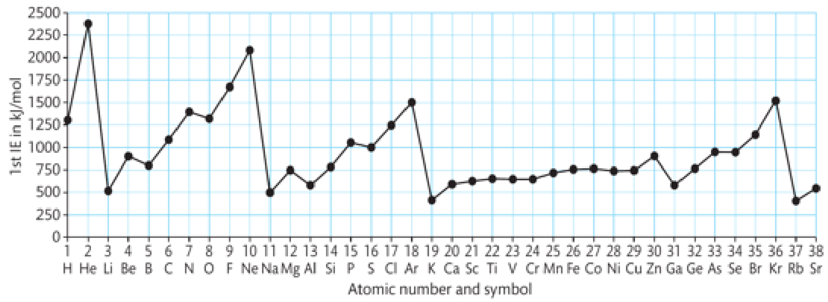


Reactivity series

Potassium	Most reactive	K
Sodium		Na
Calcium		Ca
Magnesium		Mg
Aluminium		Al
Carbon		C
Zinc		Zn
Iron		Fe
Tin		Sn
Lead		Pb
Hydrogen		H
Copper		Cu
Silver		Ag
Gold		Au
Platinum	Least reactive	Pt

First ionisation energy

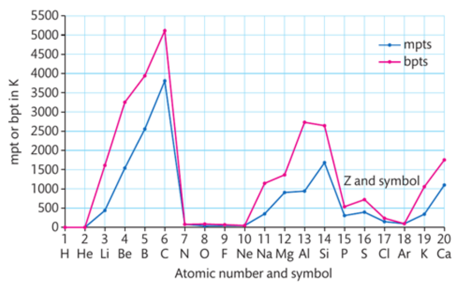
The **minimum energy** needed for one mole of **outer** electrons to be removed from one mole of gaseous atom. One more of **positively** charged ions is **formed**. First ionisation energies of elements in a period show **periodicity** (the repeating pattern seen by the elements in the periodic table).



Type of bonding

Period 2	Li	Be	B	C	N_2	O_2	F_2	Ne
Period 3	Na	Mg	Al	Si	P_4	S_8	Cl_2	Ar
Structure	giant metallic			giant covalent	simple molecular			
Forces	strong forces between positive ions and negative delocalised electrons			strong forces between atoms	weak intermolecular forces between molecules			
Bonding	metallic bonding			covalent	covalent bonding within molecules intermolecular bonding between molecules			

Melting and boiling points



Electron affinity

The change in energy when one more of a **gaseous** atom gains one more of electrons to form a mole of **negative ion**.

Element	First electron affinity kJ mol^{-1}
Fluorine	-328
Chlorine	-349
Bromine	-324
Iodine	-295

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Products and reactivity of period 2 and 3 elemenets with oxygen

Group	Element	Reactions with oxygen	Equations
1	Lithium Sodium	Rapid, burns with red flame. Metal oxide produced that forms a **alkaline solution** when dissolved in water. Very vigorous, burns with orange flame. Metal oxide produced that form basic solution when dissolved in water.	$4Li(s) + O_2(g) \rightarrow 2Li_2O(s)$ $4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$ $2Na(s) + O_2(g) \rightarrow Na_2O_2(s)$
2	Beryllium and magnesium	Needs heat to react as do group 1 elements. Very vigorous reactions.	$2Be(s) + O_2(g) \rightarrow 2BeO(s)$
3	Aluminium	Vigorous at first. Rapidly forms a water insoluble coating of Al_2O_3. This layer prevents the aluminium below from corroding and so makes aluminium an extremely useful material.	$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$ It is **amphoteric**.
4	Carbon Silicon	Forms slightly acidic oxides. Shows reaction with heat. No reaction	$C(s) + O_2(g) \rightarrow CO_2(g)$ $2C(g) + O_2(g) \rightarrow 2CO(g)$ - this is incomplete combustion $Si(s) + O_2(g) \rightarrow SiO_2(s)$ - weak acidic
5	Nitrogen Phosphorus	Forms a range of oxides with different **oxidation** states. A high temperature is needed for these reactions to take place. Burns vigorously with a white flame.	It can produce NO, and NO_2 and N_2O_5. P_4O_6 if limited oxygen, P_4O_{10} if excess oxygen.
6	Oxygen Sulfur	In ozone layer. O_2 and O_3 are **allotropes**. Two oxides form. Burns slowly with a blue flame.	$O(g) + O_2(g) \rightarrow O_3(g)$ $S(g) + O_2(g) \rightarrow SO_2(g)$ $2SO_2(g) + \frac{1}{2}O_2(g) \rightarrow 2SO_3(g)$
7	Most halogens react	Unstable oxides form.	Not usually formed by direct reaction.
0	Neon Argon	No reaction.	

Displacement reactions

A metal will **displace** a less reactive metal in a **metal salt solution**. You can predict which metals will displace which from their salts.

	Magnesium	Zinc	Iron	Copper
Magnesium sulfate	No reaction	No reaction	No reaction	No reaction
Zinc sulfate	Displacement	No reaction	No reaction	No reaction
Iron sulfate	Displacement	Displacement	No reaction	No reaction
Copper sulfate	Displacement	Displacement	Displacement	No reaction

Halogens are **oxidising agents** (withdraw electrons from other atoms/ions). The oxidising power of halogens **decreases** as you go down group 7.

	Chlorine	Bromine	Iodine
Potassium chloride	No reaction	No reaction	No reaction
Potassium bromide	Displacement	No reaction	No reaction
Potassium iodide	Displacement	Displacement	No reaction

Variable oxidation states

Transition metals have variable oxidation states due to their highest energy electrons being in the d sub-shell. When a transition metal loses electrons to form a positive ion, the 4s electrons are lost **first**, followed by the 3d electrons. The maximum oxidation state **increases** as you go along the period.

Many transition metals are used as **catalysts** (substances that increase the rate of chemical reaction but are unchanged at the end of the reaction.) E.g. iron is used in the **Haber** process and platinum is used in **catalytic converters** in cars.

Uses and applications of substances

- Metal and non-metal oxides have a range of applications. E.g. Magnesium oxide is used as a starter material for producing fibreglass.
- Metal salts are used to make colours in fireworks.
- Sodium chloride is used for many different manufacturing processes such as making glass, paper and rubber.
- Sulfates are used in detergents.
- Copper sulfate is used in water treatment to kill algae.

Products and reactivity of metals

Oxygen

Group 2 metals burn in oxygen to form metal oxides:
$2M + O_2 > 2MO$

Group 3 metals react with oxygen: $4M + 3O_2 > 2M_2O_3$

Group 4 metals can also produce oxides with the formula MO and MO_2

Water

Group 1 metals (alkali metals) produce a basic solution:
$2M(s) + 2H_2O(l) > 2M^+(aq) + 2OH^-(aq) + H_2(g)$

Group 2 metals produce hydroxides:
$M(s) + 2H_2O(l) > M(OH)_2(aq) + H_2(g)$

Group 3 metals are not very reactive with water.

Dilute acids

$Mg + 2HCl \rightarrow MgCl_2 + H_2$

It reacts with dilute sulfuric acid to give magnesium sulfate and hydrogen:

$Mg + H_2SO_4 \rightarrow MgSO_4 + H_2$

Sodium reacts with hydrochloric acid to form sodium chloride and hydrogen:

$2Na + 2HCl \rightarrow 2NaCl + H_2$

Oxidation and reduction

Redox: the transfer of electrons during chemical reactions.

Reduction: when an atom/ion gains electrons. **OIL RIG** (**O**xidation **I**s **L**oss, **R**eduction **I**s **G**ain)

Half equation: An equation that shows the loss or gain of electrons during a reaction.

Oxidation state: The number assigned to an element in a chemical compound. Also called oxidation number.

$$\text{Oxidant} + e^- \xrightarrow{\text{Reduction}} \text{Product}$$

(Electrons **gained**; oxidation number **decreases**)

$$\text{Reductant} \xrightarrow{\text{Oxidation}} \text{Product} + e^-$$

(Electrons **lost**; oxidation number **increases**)

$$\frac{1}{2}O_2 + 2e^- \rightarrow O^{2-}.$$

$$Mg \rightarrow Mg^{2+} + 2e^-$$

$$Mg + \frac{1}{2}O_2 \rightarrow MgO$$