Adrian's Study Kit

Year 11 Chemistry

buzz buzz… i'm an electron

Syllabus:
https://educationstandards.nsw.edu.au/wps/portal/nesa/11-12/stage-6-learning-areas/stage-6-science/chemistry-2017

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Properties and Structure of Matter
Module 1

matter matters!

Properties of Matter

Purity of substances

Pure substances • Elements • Compounds	• **Fixed composition** (*i.e* the proportions of atoms remain the same throughout) (*e.g* `NaCl` will always have a ratio of `1Na:1Cl`) • **Fixed properties** (*e.g* $H_2$0 will always have a boiling point of 100°C) • **Chemically bonded** (*e.g* `NaCl` is an ionically-bonded molecule) • **Can only be separated through chemical reactions** (*e.g* combustion, electrolysis)
Impure substances • Mixtures	• **Variable composition** (*e.g* salt water has $H_2$0 and `NaCl` mixed randomly) You can vary the amount of each substance in a mixture • **Variable properties** (*i.e* each substance in the mixture keeps its own properties) • **Not chemically joined** • **Can be separated by physical separation technique** (*e.g* filtration, boiling)

Elements, compounds & mixtures

	Description	**Example**
Elements	• Pure substance • Only one type of atom • Cannot be broken down into a simpler substance	• H_2 • Fe • Ar
Compounds	• Pure substance • Has multiple types of atoms • Can only be separated through chemical equations	• NaCl • H_2SO_4 • CCl_4
Mixtures	• Impure substance • Variable composition and properties • Not chemically joined	• Salt water • Air • Milk

Homogeneous and heterogeneous mixtures

Homogenous mixture	Uniform in composition	*e.g* salt water is homogeneous since the salt is uniformly mixed within the water particles.
Heterogeneous mixture	Variable composition	*e.g* muddy water is heterogeneous since heavier sediment sinks to the bottom, while the more watery mud sits at the top.

Note: Milk is heterogeneous since there are fat lumps that are distributed unevenly… yuck.

Solutions

- A homogenous mixture in which dispersed particles are so small they never settle out
- Cannot be seen by a microscope

Aqueous solutions

- Formed when a substance is dissolved in water
- The substance that is dissolved is called the **solute0**
- The liquid that dissolves the solute is called the **solvent**

Example: When dissolving salt in water, salt is the **solute** and water is the **solvent.**

Polyatomic ions

Name	Formula	Valency
Hydroxide	OH^-	1–
Nitrate	NO_3^-	1–
Hydrogencarbonate	HCO_3^-	1-
Carbonate	CO_3^{2-}	2-
Sulfate	SO_4^{2-}	2-
Phosphate	PO_4^{3-}	3-

Naming compounds

Polyatomic ions

nitr–**ate**	Largest number of oxygen atoms	NO_3
nitr–**ite**	Least number of oxygen atoms	NO_2

nitr–ide	Negatively charged ion	N_2

Ionic compounds (metal/non-metal)

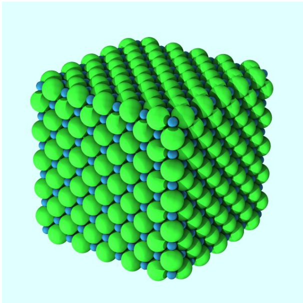
- **Magnesium chloride** — **$MgCl_2$**
- **Calcium hydroxide** — **$Ca(OH)_2$**

Covalent compounds (non-metal/non-metal)

- **Dinitrogen pentoxide** — **N_2O_5**
- **(Di)phosphorus trioxide** — **P_2O_3**

1	2	3	4	5	6	7	8	9	10
mono	di	tri	tetra	pent	hex	hepta	octa	nona	deca

Empirical formula



This is our friend, **NaCl**.

NaCl is in a 1:1 ratio. It is the simplest form in which it exists. Hence, this is known as the **empirical formula**.

Therefore, the empirical formula is the simplest ratio in which atoms in a compound can exist.

Ionic equations

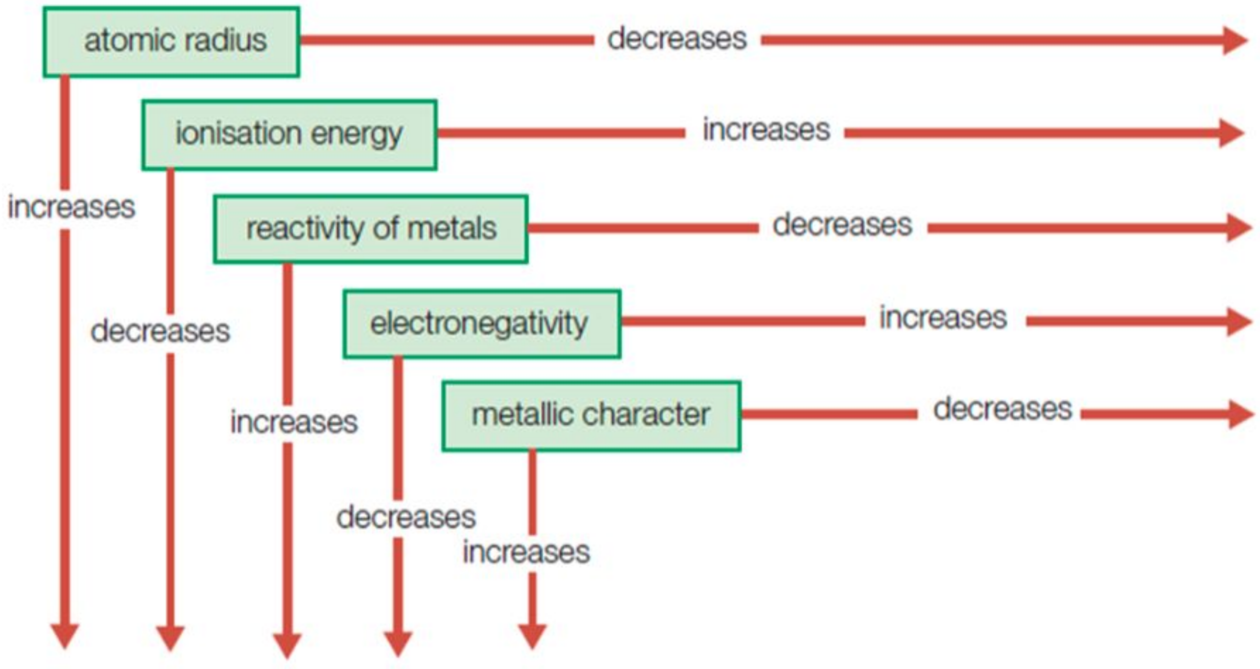
Given an element, an electron can be added to either side of an ionic equation. This will cause it to either gain/lose electrons, giving it a full valence shell.

$$Na \rightarrow Na^{+} + e^{-}$$
$$(2,\ 8,\ 1) \rightarrow (2,\ 8) + e^{-}$$

$$F + e^{-} \rightarrow F^{-}$$
$$(2,\ 7) + e^{-} \rightarrow (2,\ 8)$$

Periodicity

There are many other trends that we can identify throughout the Periodic Table. It can be summarised in the following diagram.



let's try and do our best to understand each one!

Atomic radius

Atomic radius is how large an atom is. Yep, that's about it.

Trends in atomic radius

Excluding the noble gases, atomic radius:

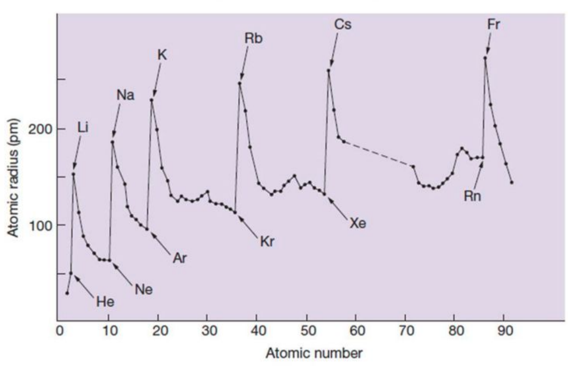
- **<u>decreases</u>** across a period from left to right
- **<u>increases</u>** down a group from top to bottom

Note: Note that along a period, you gain more protons, yet the atomic radius decreases. Therefore, it can be said that **as atomic radius decreases, atomic density increases**.

Atomic radius will suddenly increase at the beginning of every period. This occurs as at the beginning of each period, it gains a new electron shell.

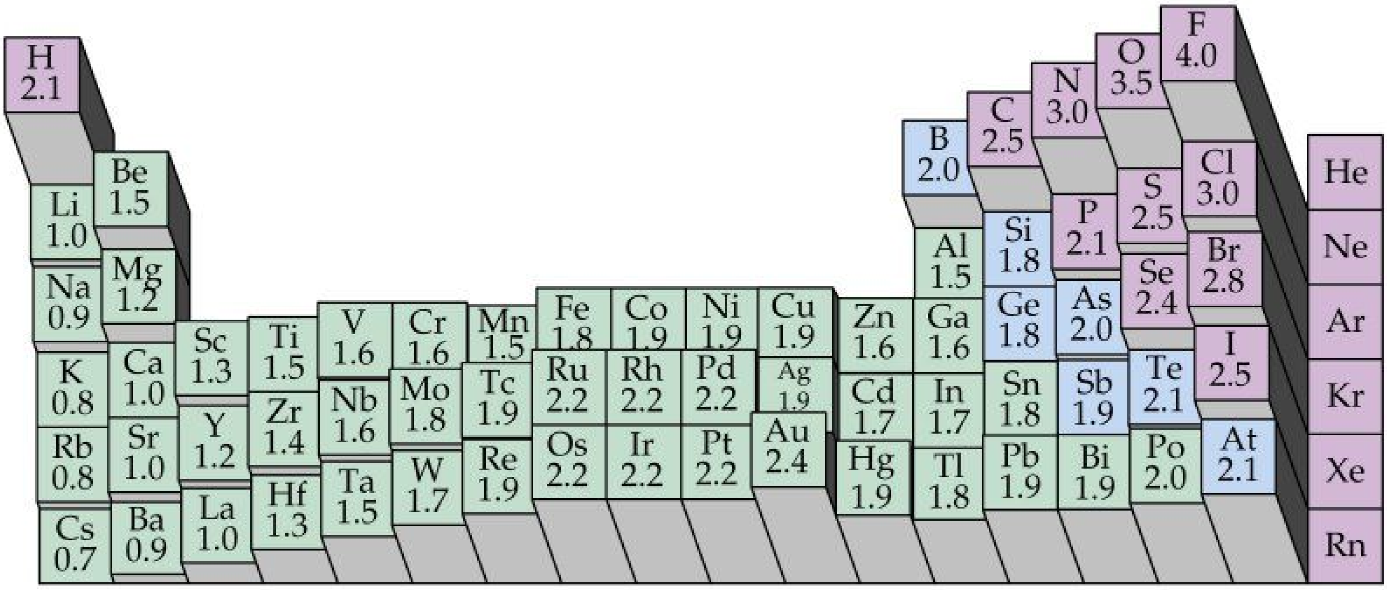
Why does atomic radius decrease across a period and increase down a group? See [Explaining core charge](#).

Atomic radius increases in going down any group of the table.



Electronegativity

Electronegativity is the measure of an atom's ability to attract shared electrons towards itself, when forming compounds.



Explaining core charge

Electronegativity and atomic radius is best explained with the concept of **core charge:**

1. At the beginning of each period, the **core charge** of an atom suddenly increases, given the addition of one electron shell.
2. Down a group, the core charge remains **constant**, however valence electrons have weaker attractions as they're further away from the nucleus.
3. Across a period, core charge increases ∴ the valence electrons are attracted more towards the nucleus.

The **core charge** of an atom is a measure of the attraction of valence electrons toward the nucleus.

- **Down a group,** atoms will contain the same core charge. But, since they're further away, the valence electrons are held less strongly as the force of attraction is not as strong.
- **Across a period**, this core charge increases. Therefore, valence electrons are attracted more strongly towards the nucleus.

Therefore, core charge can be calculated using the following formula:

$$core\ charge\ =\ number\ of\ p^{+}\ -\ number\ of\ inner\ shell\ e^{-}$$

Example:
Take two Group 7 non-metals.

- **Fluorine** has electron configuration (2, 7) = $9p^{+}$
  ∴ inner-shell e^{-} = 2
  ∴ core charge = 7+
- **Chlorine** has electron configuration (2, 8, 7) = $17p^{+}$

∴ inner-shell e^- = 10
∴ core charge = 7+

Fluorine and chlorine carry the same core charge of 7+. However, fluorine's valence electrons are closer to the nucleus, therefore it will carry a stronger force of attraction.

Trends in electronegativity

Excluding the noble gases, electronegativity and ionic character:

- **increases** across a period from left to right
- **decreases** down a group from top to bottom

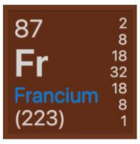
Since ionic character increases across a period, the type of compound will demonstrate trends as well. For instance, take a compound with oxygen, using various elements across period 3.

Across a period

Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_3	Cl_2O_7
Ionic			Covalent		Covalent/ionic	

Down a group

CH_4	Covalent
$SiCl_4$	
$SnCl_4$	Ionic
PbO	



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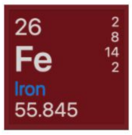

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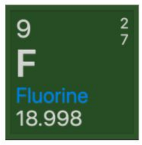


3.8

Note: Percent ionic character is the amount of electron sharing between two atoms. A percent ionic character is proportional to Δ electronegativity.

If $\Delta > 2.1$ ∴ ionic compound, otherwise covalent.

Note:

- Electro**negativity** = tendency to **gain** electrons.
- Electro**positivity** = tendency to **lose** electrons

Example:

- Fluorine has the highest electronegativity of any element on the Periodic Table. This means that it attracts electrons better than any other element. It also has the highest ionic character.
- Caesium has a very low electronegativity. It does not attract electrons very well.

Bonding

Revision: Ionic and covalent bonds

- **Ionic bonding:** transfer of electrons between ions
- **Covalent bonding:** sharing of electrons

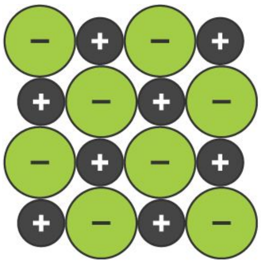
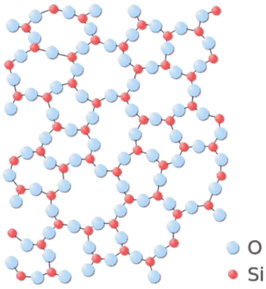
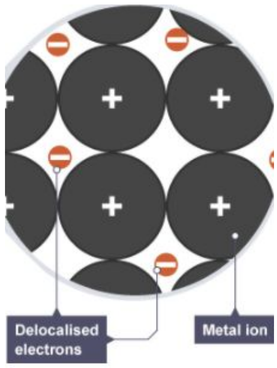
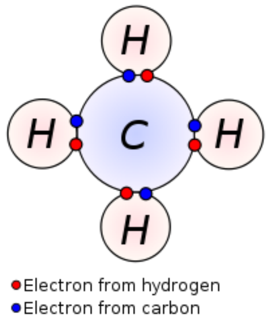
What is intramolecular bonding?

Intramolecular forces are forces between molecules, such as ionic bonds or covalent bonds.

Chemical structures

Summary

	Lattice			
	Ionic networks	**Covalent networks**	**Metallic structure**	**Covalent molecular**
Description	Ionic compounds which extend indefinitely in a regular, repeating arrangement.	Solids where covalent bonding extends indefinitely in a regular, repeating arrangement.	Metals where positive metal ions are arranged in a lattice structure, amongst a sea of delocalised electrons.	Sharing of electrons.
Melting point/boiling point	High (strong ionic bonds)	Very high (strong covalent bonds)	High (strong metallic bonds)	Low
Solubility	Soluble	Insoluble	Insoluble	Insoluble=
Electrical conductivity	Solid: no Molten/liquid: yes as there are free moving electrons	No	Yes, as there are free moving electrons in the sea of delocalised electrons	No
Hardness	Hard and brittle	Hard and brittle	Malleable	Soft
Forces	Electrostatic forces of attraction between oppositely charged ions	Covalent bonding throughout the crystal	Electrostatic forces of attraction between electrons and positive ions (metallic bonding)	Intermolecular forces
Examples	Sodium chloride	Diamond Silicon dioxide (sand)	Copper	Bromine gas Oxygen gas Water

Did you know? Silicon dioxide was the **molecule of the month** for November 2013, as designated by Guillermo Godino Sedano of King's College, Madrid, Spain.

Melting point/boiling point

Melting point and boiling point (mp/bp) is dependent on how much energy is required to break bonds within substances.

- **Ionic networks** have strong electrostatic forces of attraction between the positive and negative ions, and therefore will have a **high mp/bp.**
- **Covalent networks** have strong covalent bonding within the crystal, and therefore will have a **very high mp/bp.**
- **Metallic structures** have strong electrostatic forces of attractions between positive ions and the sea of delocalised electrons, and therefore will have a **high mp/bp.**
- **Covalent molecular solids** have weak covalent forces of attraction, and therefore will have a **low mp/bp.**

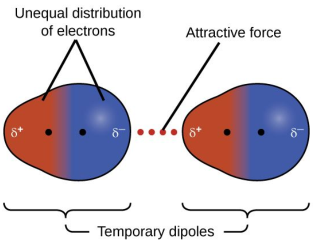
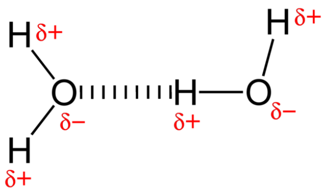
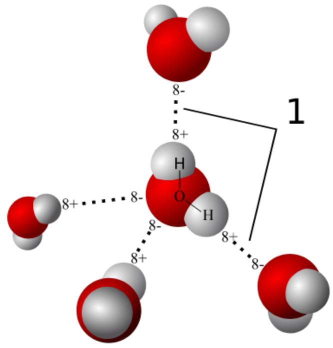
Electrical conductivity

Electrical conductivity is dependent on the number of free flowing electrons.

- **Ionic networks** do not conduct electricity as a solid as there are no free electrons. But in a molten/liquid state, it can, since the ions are free to move around.
- **Covalent networks** can't conduct electricity as there are no free electrons.
- **Metallic structures** can conduct electricity through a sea of delocalised electrons.
- **Covalent molecules** can't conduct electricity as there are no free ions or electrons.

Intermolecular forces

Intermolecular forces are forces between molecules that attract them together.

	Dispersion forces	Dipole-dipole forces	Hydrogen bonds
Between which molecules?	All atoms and molecules	Polar molecules only	Hydrogen attached to fluorine, oxygen or nitrogen.
Explanation	Constant and fluctuating movement of electrons and the nucleus, causes constant repulsion and attraction.	Permanent dipoles between a slightly positive side and slightly negative side of each respective molecule.	A form of dipole-dipole force between bonds with very high electronegativity (H-F, H-O, H-N).
Diagram			
Increases with what?	Increases with higher molecular mass.	Increases with difference in electronegativity (aka dipole moment) which is representative of polarity.	
Relative strength	Weakest	Middle	Strongest
Example	H_2O, CO, Cl_2	CCl_4, HCl, H_2O	HF, H_2O

Note: You can remember hydrogen bonds as H attaching to only **F, O, N**.

References and Related Links

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- **Periodicity,** BBC Bitesize (https://www.bbc.co.uk/bitesize/guides/zxc99j6/revision/1)
- **Module 1** on Canvas (https://staloysius.instructure.com/courses/3932/files/folder/Module%201)
- **Modelling metals,** BBC Bitesize (https://www.bbc.co.uk/bitesize/guides/z9m6v9q/revision/1)
- **Molecule of the Month** (http://www.chm.bris.ac.uk/motm/motm.htm)
- **Year 11 Chemistry Depth Study** (https://docs.google.com/presentation/d/174Nier8TmE6O_hm6bY1_vKWes8I3myPnku74lbiHQck/edit?usp=sharing)

Image credits

- **Ionic network:** BBC Bitesize (https://www.bbc.co.uk/bitesize/guides/ztc6w6f/revision/3)
- **Metallic structure:** LibreTexts (https://chem.libretexts.org/@go/page/1989)
- **Covalent networks:** BBC Bitesize (https://www.bbc.co.uk/bitesize/guides/zxc99j6/revision/4)
- **Covalent molecular:** Dynablast (https://commons.wikimedia.org/wiki/File:Covalent.svg)

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