



chemistry

acids and bases (incl. air and atmosphere)

acids

▼ definition

An **acid** is a substance that **dissociates** or **ionises** to form **H⁺** ions when dissolved in water.

▼ properties

1. **sour taste**— citrus fruits like lemons and oranges contain citric acid, and taste sour
2. **turns blue litmus paper red**

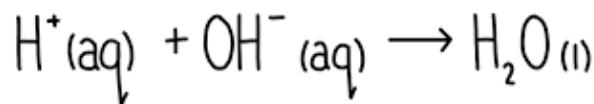
▼ reactions

1. acid + **carbonate/ hydrogen carbonate**—> salt + water + carbon dioxide
2. acid + **reactive metal** —> salt + hydrogen
3. acid + **metal oxide/hydroxide/aqueous ammonia**—> salt + water (**neutralisation**)
4. acid + **sodium sulfite** (SO₃ 2-) —> sodium salt + sulfur dioxide + water

However, not all metals will react with acids to give a salt and hydrogen gas. When **unreactive metals**, such as copper, silver, or gold are added to dilute acids, there is no reaction.

Lead appears not to react with dilute hydrochloric acid and dilute sulfuric acid. The initial reaction between lead and dilute acid produces a layer of **lead(II) chloride or lead(II) sulfate** which is **insoluble in water** and quickly **forms a coating around the metal**. The **coating prevents the metal from further reaction with the acid**.

ionic equation of acid + alkali:



▼ tests for gases

1. hydrogen gas: place a **lighted** splint **at the mouth of the test tube** when effervescence is observed. Hydrogen gas will **extinguish** the lighted split with a ‘**pop**’ sound
2. carbon dioxide gas: Bubble the gas through **limewater** (aqueous calcium hydroxide). Carbon dioxide forms a **white precipitate** with limewater.
3. sulfur dioxide gas: place a strip of filter paper soaked in **acidified potassium manganate (VII)** solution in the gas sample. Sulfur dioxide gas turns **purple** acidified potassium manganate (VII) **colourless**. Sulfur dioxide also turns **orange** acidified potassium dichromate (VI) **green**.
4. ammonia gas: place a moist red litmus paper in the gas sample. Ammonia gas turns moist red litmus paper **blue**. Ammonia gas is **colourless** and **pungent**.

▼ strength

The strength of an acid or base refers to the extent to which an acid or base dissociates when dissolved in water.

A **strong acid** is defined as an acid that **dissociates (or ionises) completely in water** to form hydrogen ions, H⁺

A **weak acid** is defined as an acid that **dissociates (or ionises) partially in water** to form hydrogen ions, H⁺

An **electrolyte** is a substance that **conducts an electric current** as a result of its dissociation to form ions when dissolved in a solvent such as water. There is a presence of freely moving/mobile ions that act as **mobile charge carriers** that can conduct electricity.

A **strong electrolyte** completely dissociates in solution and contains only ions. Strong electrolytes are good conductors of electricity (eg. strong acids and strong bases)

▼ concentration

the amount of substance (solute) dissolved in a fixed volume of solution

- the strength of an acid or base cannot be changed
- the concentration of an acid or base can be changed.

bases

▼ definition

A base is a substance that **reacts with an acid** to produce **salt** and **water** only. All **metal oxides** and **hydroxides** can act as bases

Bases: all metal oxides and hydroxides + aqueous ammonia

Soluble bases are called alkalis

▼ properties of alkali

1. bitter taste and soapy feel
2. turn red litmus paper blue

▼ reaction of alkalis

1. alkali + ammonium (NH_4) salt \rightarrow salt + water + ammonia gas
2. alkali + acid \rightarrow salt + water

▼ strength of alkalis

The strength of an acid or base refers to the extent to which an acid or base dissociates when dissolved in water.

A **strong base** is defined as an acid that **dissociates (or ionises) completely in water** to form hydroxide ions, OH^- .

A **weak acid** is defined as an acid that **dissociates (or ionises) partially in water** to form hydroxide ions, OH^- .

An **electrolyte** is a substance that **conducts an electric current** as a result of its dissociation to form ions when dissolved in a solvent such as water. There is a presence of freely moving/mobile ions that act as **mobile charge carriers** that can conduct electricity.

A **strong electrolyte** completely dissociates in solution and contains only ions. Strong electrolytes are good conductors of electricity (eg. strong acids and strong bases)

▼ concentration of alkalis

the amount of substance (solute) dissolved in a fixed volume of solution

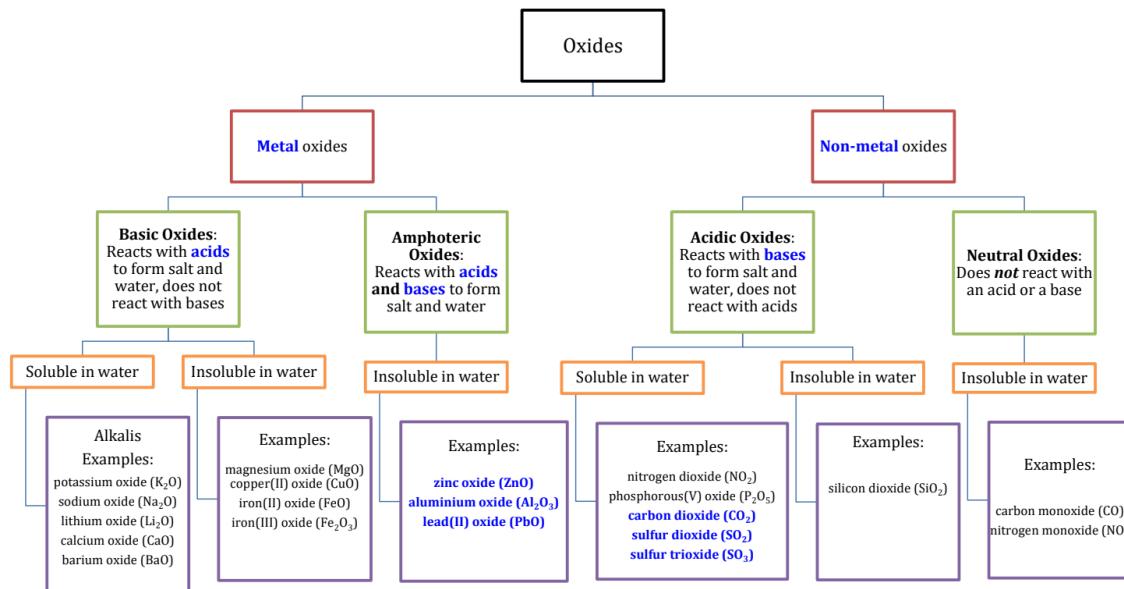
- the strength of an acid or base cannot be changed
- the concentration of an acid or base can be changed.

▼ pH scale and indicators

pH Indicator	Colour		pH range at which indicator changes colour
	Acidic	Basic	
litmus	red	blue	5 – 8
methyl orange	red	yellow	3 – 5
screened methyl orange	pink	green	3 – 5
thymol blue	yellow	blue	8 – 10
phenolphthalein	colourless	pink	8 – 10

pH	0 – 1	2 – 4	5 – 6	7	8 – 10	11 – 12	13 – 14
Colour	red	orange	yellow	green	dark green	blue	violet

▼ classification of oxides and their reactions



▼ effects of air pollution

Air pollutant	Effects
carbon monoxide (CO)	<ul style="list-style-type: none"> reacts with haemoglobin in blood to form carboxyhaemoglobin, which reduces the ability of haemoglobin to transport oxygen to the rest of the body causes headaches, fatigue, breathing difficulties and even death
sulfur dioxide (SO ₂)	
oxides of nitrogen (NO, NO ₂)	<ul style="list-style-type: none"> irritate the eyes and lungs and cause breathing difficulties high levels lead to inflammation of the lungs (bronchitis) react with water in the atmosphere to form acid rain, which corrodes buildings and harms aquatic life and plants
unburnt hydrocarbons	<ul style="list-style-type: none"> cause cancer (carcinogenic) react with nitrogen oxides to form ozone
methane	<ul style="list-style-type: none"> is a greenhouse gas, which causes global warming
ozone	<ul style="list-style-type: none"> forms photochemical smog, which irritates the eyes and lungs and causes breathing difficulties damages crops

acid rain:

formed when acidic air pollutants such as sulfur dioxide and nitrogen dioxide react with water in the atmosphere.

sulfur dioxide — dissolved in water and then oxidized to form sulfuric acid

nitrogen dioxide — converted to nitric acid in the presence of oxygen and water.

Effects of acid rain on	Damages
buildings	Acid rain reacts with metals and carbonates like marble and limestone. When this happens, metal bridges and stone buildings are damaged.
plants Al(OH)_3	Acid rain leaches nutrients from soil and causes plants to wither and die. Acid rain may also react with aluminum hydroxide (Al_2O_3) in the soil to produce Al^{3+} ions which are toxic to plants.
aquatic life	Acid rain reduces the pH value of natural water bodies like lakes and streams, from between 6.5 and 8.5 to below 4, which is too acidic for fish and other aquatic life to survive.

sample questions

- ▼ compare solutions of 0.100mol/dm³ hydrochloric acid and 0.100mol/dm³ of sulfuric acid. Which will have a higher concentration of H⁺ ions? explain your choice.

Sulfuric acid will have a higher concentration of H⁺ ions. The concentration and strength of both acids are the same, i.e. strong acids. However, their basicity is different. Sulfuric acid is a dibasic acid, while hydrochloric acid is a monobasic acid. Hence, sulfuric acid will produce twice,double of H⁺ ions, while hydrochloric acid will produce 0.100mol/dm³ of H⁺ ions.

- ▼ State and explain the difference between the electrical conductivity of a strong acid and a weak acid of the same concentration.

For the same concentration, a strong acid has a higher electrical conductivity than a weak acid. A strong acid undergoes complete dissociation in water to form ions which gives rise to a higher concentration of mobile ions to act as mobile charge carriers compared to weak acids, which undergo partial dissociation in water to form a lower concentration of mobile ions.

- ▼ If all the hydrochloric acid was removed, how do you think this would affect the test results? explain.

There will be no visible reaction for each test. In the absence of water, hydrochloric acid would not dissociate to form hydrogen ions. Hydrogen ions are responsible for the chemical properties of acids. The lack of hydrogen ions means there is no visible reaction for each test.

- ▼ Hydrogen chloride gas, HCl, was bubbled and dissolved in ethanol, an organic solvent. A piece of magnesium ribbon is then added to this solution. State what you would expect to observe.

no visible reaction

▼ what can you infer about the role of solvent in acidity?

hydrogen chloride is a covalent compound. In an organic solvent, hydrogen chloride exists as molecules. Hence it does not behave as an acid. However, in an aqueous solvent, hydrogen chloride dissociates to form hydrogen ions which are responsible for the acidic properties.

▼ **What is the limitation of using methyl orange compared to a universal indicator?**

There is no way we can determine the exact pH of the substance OR whether the substance is neutral or alkaline. If methyl orange turns yellow, we can predict that the pH of substance is 4.0 and if methyl orange turns red, we can predict that the pH is below 3.0. We can use the various distinctive colour changes of the universal indicator paper to estimate the exact pH of the substance and determine if its acidic, neutral or alkaline.

▼ **suggest why universal indicator is not usually used in titrations**

the colour change in universal indicator is gradual giving different colour for different pH values [1] hence the end point cannot be easily determined. [1].

▼ **describe an experiment (using materials other than pH sensor or Universal Indicator) to show that ethanoic acid is a weak acid and hydrochloric acid is a strong acid.**

Use an electrical conductivity probe to measure the electrical conductivity of both acids. HCl, a strong acid, will give a greater electrical conductivity value than ethanoic acid, a weak acid.

Connect 2 graphite rods to a circuit with a light bulb. Dip both ends of the graphite rods into each of the acid. The bulb lights up brightly in HCl as it is a strong acid which the bulb lights dimly in ethanoic acid.

To a portion of each acid, add a piece of magnesium ribbon or a spatula od calcium carbonate. A lot of effervescence per unit time is observed for the strong acid while little or no effervescence is observed for the weak acid.

▼ **Why is it important not to add ammonium compounds to soils that have been treated with calcium hydroxide?**

This is because the ammonium compounds will react with calcium hydroxide resulting in a loss of nitrogen in the form of ammonia gas.

salts and ionic equations

▼ **salts definition**

A substance formed when a metallic ion or an **ammonium ion (NH_4^+) replaces one or more hydrogen ions of an acid**

It is an ionic compound that consists of a cation other than H⁺ and an anion other than OH⁻ or O₂⁻

▼ solubility table

Soluble in water	Insoluble in water
All Group I and ammonium salts are soluble	
All nitrates are soluble	
All sulfates are soluble except	PbSO ₄ , BaSO ₄ Note that CaSO ₄ and Ag ₂ SO ₄ are sparingly soluble
All chlorides/bromides/iodides are soluble except	silver and lead(II) chloride/iodide/bromide
except Group I and ammonium carbonates e.g. Na ₂ CO ₃ , K ₂ CO ₃ , (NH ₄) ₂ CO ₃	All carbonates are insoluble
except Group I oxides e.g. Na ₂ O, K ₂ O	All metal oxides are insoluble Note that oxides of Group II elements (Ca, Sr, and Ba) increases in solubility from calcium down. CaO is sparingly soluble. Solubility of SrO and BaO increases in water.
except Group I hydroxides e.g. NaOH, KOH	All hydroxides are insoluble Note that hydroxides of Group II elements (Ca, Sr, and Ba) increases in solubility from calcium down. Ca(OH) ₂ is sparingly soluble and the solubility of Sr(OH) ₂ and Ba(OH) ₂ increases in water.

preparation of salts

▼ Precipitation

Products: insoluble

Reactants: 2 soluble (aqueous)

Procedure:

1. Mix ___ solution and ___ solution together
2. Stir constantly to ensure the mixture is homogenous
3. Filter the mixture and collect the residue. The residue is ___
4. Wash the ___ precipitate with distilled water.
5. Dry the ___ residue between pieces of filter paper.

▼ Titration

Products: Soluble (NH_4^+ or Group 1)

Reactants: 2 Soluble (aqueous) — involving acid + soluble carbonate/alkali

Procedure:

1. Fill the burette with dilute ___ acid.
2. Pipette 25.0cm³ of hydroxide/carbonate solution into a conical flask
3. Add a few drops of methyl orange indicator to the solution in the conical flask
4. Titrate the hydroxide/carbonate solution with the dilute hydrochloric acid until the indicator changes from red to orange.
5. Note the volume of dilute hydrochloric acid used
6. Repeat the experiment without the indicator
7. Add the same amount of dilute hydrochloric acid into the burette. This gives pure ___ solution, without excess acid or alkali
8. Heat the ___ solution gently until it is saturated
9. Allow the saturated solution to cool until crystals form
10. Filter the mixture to obtain the ___ crystals
11. Wash the crystals with a small amount of cold, distilled water
12. Dry the crystals between sheets of filter paper.

▼ Reaction of insoluble substance with acid (AEISTA)

products: soluble

reactants: 1 soluble and 1 insoluble (aqueous + solid)

acid + reactive metals + insoluble base + insoluble metal carbonate

do NOT use highly reactive metals (group 1 metals) and unreactive metals (copper, silver, gold)

Procedure:

1. Fill half a beaker with warm ____ acid
2. With constant stirring, add the solid oxide/carbonate to the acid with stirring until the solid is added in excess or no more effervescence is observed
3. Filter the excess unreacted solid ____ oxide/carbonate and collect the filtrate. The filtrate is ____ solution.
4. Heat the solution gently until it is saturated
5. Allow the saturated solution to cool for the hydrated ____ crystals to form.
6. Filter the mixture to collect the hydrated ____ crystals
7. Wash the crystals with cold, distilled water.
8. Dry the crystals between pieces of filter paper.

writing of ionic equations

▼ steps

1. write down the **balanced chemical equation** of the reaction, complete with state symbols
2. **Write the chemical reaction in terms of ions for aqueous state only.** Substances in solid, liquid and gaseous state do not dissociate.
3. Cancel out **spectator ions**
4. Write the ionic equation by **leaving out the spectator ions and reducing to the simplest ratio.**

sample questions

▼ excess aqueous barium nitrate is added to aqueous aluminium sulfate, and the precipitate is removed by filtration. What are the main ions in the filtrate?

barium sulfate: precipitate

aluminium nitrate solution: filtrate

barium nitrate is added in excess, barium ions will be present.

▼ With reference to the method of titration, predict the solubility of the salt being prepared and why this method is preferable for producing this type of salt.

the salt is soluble in water. As the starting reagents are also soluble in water, using an indicator allows the exact volumes of either reactants needed for complete reaction to be determined. With no excess of either reactant.

Hence a pure salt can be obtained when we repeat the titration without an indicator.

▼ A student was given a sample of lead(II) carbonate. Describe how he can use it to prepare some dry lead(II) sulfate in the laboratory.

1. Add excess lead(II) carbonate to nitric acid in a beaker. — (cannot directly add lead(II) carbonate to sulfuric acid as a coating of lead(II) sulfate will form and stop the reaction)
2. Filter to remove the unreacted lead(II) carbonate.
3. Add aqueous ammonium sulfate/sodium sulfate/potassium sulfate/ dilute sulfuric acid to the filtrate (lead(II) nitrate)
4. A precipitate of lead (II) sulfate forms
5. Filter to obtain the precipitate and dry it between 2 pieces of filter paper.

Mole Concept and Stoichiometry

definitions

▼ relative atomic mass

The relative atomic mass (Ar) is the '**average**' **mass of an atom of the element** (taking into account the abundances of naturally occurring isotopes of the element) when compared to the mass 1/12 of an atom of 12C.

▼ relative molecular mass

mass of one molecule of a **covalent substance** when compared to the mass of 1/12 of an atom of carbon-12, 12C.

The relative molecular mass of any substance is the **sum of the relative atomic masses of the constituent elements** present in a molecule of that substance.

▼ relative formula mass

mass of one formula unit of an **ionic compound** when compared to the mass of 1/12 of an atom of carbon-12, 12C.

The relative formula mass of an ionic compound is obtained by **adding the relative atomic masses of the constituent elements** present in the formula unit.

▼ mole

A mole is the **amount of substance** that contains the same number of particles as the **number of atoms in exactly 12g of carbon-12 isotope**.

Avogadro's number: 6.02×10^{23}

▼ molar mass

The molar mass of a substance is the sum of its relative atomic masses of all the elements that form the substance expressed in grams per mole.

▼ molar gas volume

Avogadro's Law states that **equal volumes of all gases**, at the same temperature and pressure, **contain the same number of molecules** (or atoms if the gas is monoatomic)

▼ percentage composition

$$\text{Percentage by mass of an element in a compound} = \frac{A_r \text{ of element} \times \text{number of atoms in formula}}{M_r \text{ of compound}} \times 100\%$$

▼ relative abundance

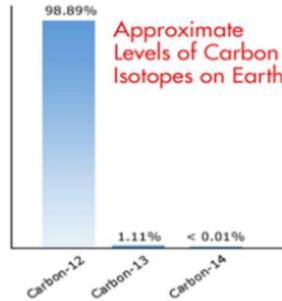
Example 1

Based on the graph on the right, what is the relative atomic mass of carbon?

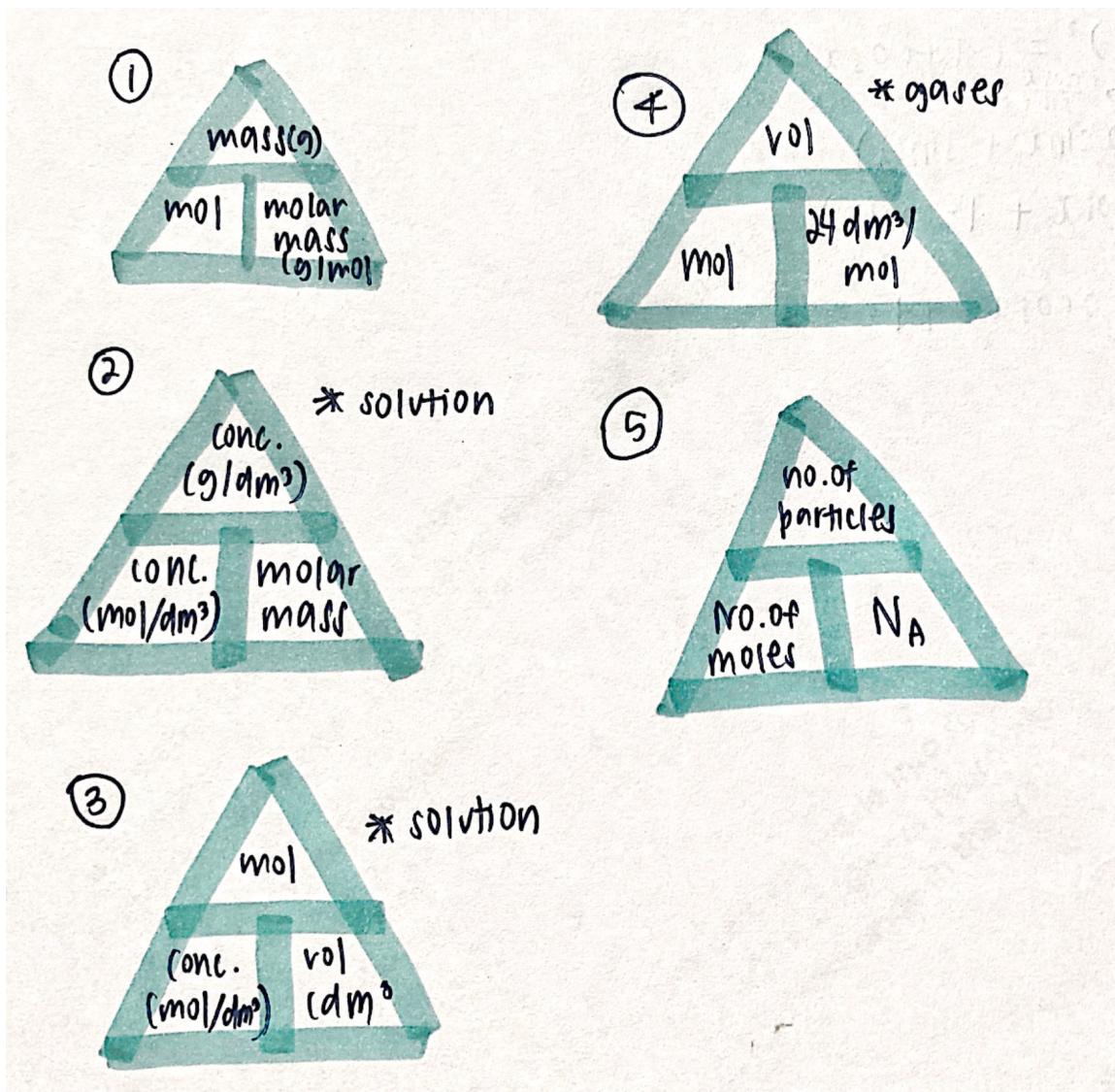
Solution:

$$A_r \text{ of carbon} = 0.9889 \times 12 + 0.0111 \times 13 = 12.1$$

(The abundance of isotope carbon-14 was not taken into account as it is less than 0.01%.)



▼ formula triangles



▼ empirical and molecular formula

Example 1: Find the empirical formula of the substance formed when 14.0g of iron combined with 6.00g of oxygen.

Solution:

Element	Iron	Oxygen
Mass / g	14.0	6.00
Relative atomic mass	56	16
No. of moles / mol	$\frac{14.0}{56} = 0.250$	$\frac{6.00}{16} = 0.375$
Mole ratio	$\frac{0.250}{0.250} = 1$	$\frac{0.375}{0.250} = 1.5$
Simplest whole number ratio	2	3

The empirical of the substance is Fe₂O₃.

$$n = \frac{\text{relative molecular mass}}{M_r \text{ from empirical formula}}$$

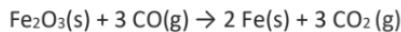
▼ stoichiometric calculations involving limiting reactants

General Approach to Solve Questions with Limiting Reagent

1. Information on at least 2 reactants will be given
 - a) as masses (g, kg or ton) [1 ton = 1000 kg = 10^6 g]
 - b) as volume of gases (cm^3 or dm^3)
 - c) as volume (cm^3 or dm^3) and concentration (mol / dm^3) of solutions
2. Find the number of moles of the 2 reactants
 - a) Convert the mass of a substance to amount (mol)
 - b) Convert the volume of gas to amount (mol)
 - c) Multiply volume (cm^3 or dm^3) x concentration (mol / dm^3) to find amount of solute (mol)
3. Compare the mole ratio to determine the limiting reagent.
4. Use the **no. of moles of the limiting reagent** to determine
 - a) Number of moles of product formed OR
 - b) Number of moles of excess reactant used.

Example 2

Iron(III) oxide (Fe_2O_3) combines with carbon monoxide to form iron metal and carbon dioxide by the reaction:



If 200 g of iron(III) oxide are reacted with 268 g of carbon monoxide,

- a) Which reactant is the limiting reagent?
- b) What is the mass of iron produced?

Solution:

$$\begin{aligned} \text{a) No. of moles of iron(III) oxide} &= 200 / (56 \times 2 + 16 \times 3) = 1.25 \text{ mol} \\ \text{No. of moles of carbon monoxide} &= 268 / (12 + 16) = 9.57 \text{ mol} \end{aligned}$$

From the balanced chemical equation, since 1 mol of iron(III) oxide reacts with 3 mol of carbon monoxide,

1.25 mol of iron(III) oxide will react completely with $1.25 \times 3 = 3.75$ mol of carbon monoxide
Hence carbon monoxide is in excess (9.57 mol), and **iron(III) oxide is the limiting reagent**.

▼ calculation of percentage yield and percentage purity .

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

$$\% \text{ purity} = \frac{\text{Mass of pure substance}}{\text{Mass of substance used in reaction}} \times 100\%$$

Atomic structure

▼ atomic (proton) number

The number of protons in an atom

▼ nucleon (mass number)

The total relative mass of an atom

▼ Isotopes

▼ definition

Isotopes are atoms of the same element but with a different number of neutrons. Most elements that commonly occur are made up of isotopes

▼ calculation of relative atomic mass

The relative atomic mass of an element is the **weighted average** of the atomic mass of the isotopes

For example, chlorine exists as two isotopes, chlorine-35, ^{35}Cl and chlorine-37, ^{37}Cl . A sample of chlorine gas consists of 75% of chlorine-35 and 25% of chlorine-37. The relative atomic mass of chlorine is calculated as follows:

$$\begin{aligned}\text{Relative atomic mass of chlorine} &= (0.75 \times 35) + (0.25 \times 37) \\ &= 35.5\end{aligned}$$

▼ writing of spdf electronic configuration

Element	Atomic number	Shorthand notation (if applicable)	spdf notation
H	1	-	$1s^1$
He	2	-	$1s^2$
Li	3	[He]2s ¹	$1s^22s^1$
Be	4	[He]2s ²	$1s^22s^2$
B	5	[He]2s ² 2p ¹	$1s^22s^22p^1$
C	6	[He]2s ² 2p ²	$1s^22s^22p^2$
N	7	[He]2s ² 2p ³	$1s^22s^22p^3$
O	8	[He]2s ² 2p ⁴	$1s^22s^22p^4$
F	9	[He]2s ² 2p ⁵	$1s^22s^22p^5$
Ne	10	[He]2s ² 2p ⁶	$1s^22s^22p^6$
Na	11	[Ne]3s ¹	$1s^22s^22p^63s^1$
Mg	12	[Ne]3s ²	$1s^22s^22p^63s^2$
Al	13	[Ne]3s ² 3p ¹	$1s^22s^22p^63s^23p^1$
Si	14	[Ne]3s ² 3p ²	$1s^22s^22p^63s^23p^2$

7

P	15	[Ne]3s ² 3p ³	$1s^22s^22p^63s^23p^3$
S	16	[Ne]3s ² 3p ⁴	$1s^22s^22p^63s^23p^4$
Cl	17	[Ne]3s ² 3p ⁵	$1s^22s^22p^63s^23p^5$
Ar	18	[Ne]3s ² 3p ⁶	$1s^22s^22p^63s^23p^6$
K	19	[Ar]4s ¹	$1s^22s^22p^63s^23p^64s^1$
Ca	20	[Ar]4s ²	$1s^22s^22p^63s^23p^64s^2$
*Cu	29	[Ar]3d ¹⁰ 4s ¹	$1s^22s^22p^63s^23p^63d^{10}4s^1$

▼ ionisation energy

▼ definition

The ionisation energy (IE) of an element is the energy required to **remove** an electron (1 mole) from its gaseous atom/ion in its ground state. It measures the **ease of an atom/ion losing an electron.**

$X(g) \rightarrow X^+(g) + e^-$	1 st ionisation energy
$X^+(g) \rightarrow X^{2+}(g) + e^-$	2 nd ionisation energy
$X^{2+}(g) \rightarrow X^{3+}(g) + e^-$	3 rd ionisation energy

sample questions

▼ explain which group of the periodic table element _ is in.

The largest increase is between the _th and _th ionisation energy. the _th electron is removed from the inner principal quantum shell while the _th electron is removed from the valence shell. Therefore there are _ electrons in the valence shell. The element is in group ____.

▼ write the outer shell electronic configuration for the element.

___ where n is the principal quantum number of the valence shell

▼ which of these elements is most likely to form an ion with a charge of 1+? Give reasons for your answer.

The largest increase is between the 1st and 2nd ionization energy. The 2nd electron is removed from the inner principal quantum shell while the 1st electron is removed from the valence shell. Therefore there is 1 electron in the valence shell, and the ion loses this electron to form an ion of 1+ charge to achieve a noble gas electronic configuration.

▼ explain the difference between the third I.E. of both elements.

The 3rd IE of X is much larger than Y. The 3rd electron is removed from an inner principal quantum shell in X while the 3rd electron is removed from the valence shell in Y. The EFOA between the 3rd electron and nucleus in X is stronger. Hence more energy is required for removal of the 3rd electron of X.

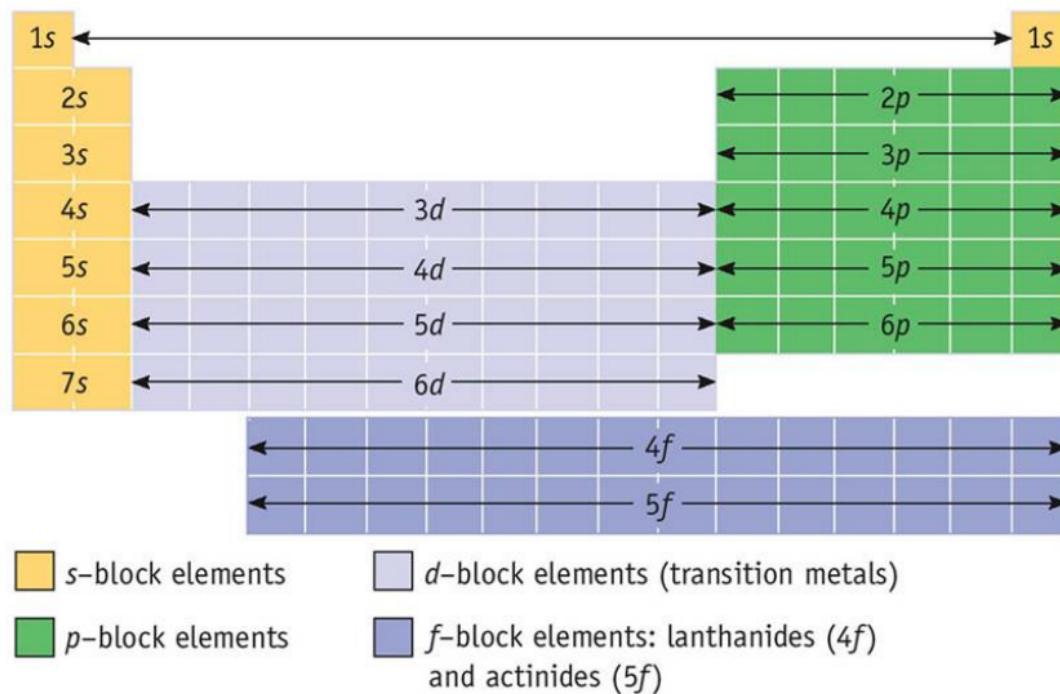
▼ The first ionisation energy of Metal X is 738 while that of metal Y is 419, what can you infer from this statement?

The valence electron of Metal X experiences a greater force of attraction from the nucleus compared to metal Y. Less energy is required to remove the valence electron from Metal Y atom. Metal Y is more reactive than Metal X.

Periodic Table

▼ Features of the periodic table

- elements are arranged according to increasing atomic number
- The **Group number** gives the number of valence electrons in the atom
- The **Period number** gives the number of principal quantum shells in the atom.
- metals → metalloids → non-metals
- basic → amphoteric → acidic



▼ Periodic trends

Atomic radius

▼ Across the Period

There is a general **decrease** in atomic radii

- Nuclear charge increases** as the number of protons increases successively from one element to the next.
- Number of inner electrons remains the same** as successive electrons are added to the same principal quantum shell
- Electrostatic forces of attraction between the valence electron and the nucleus **increases** and the size of the electron cloud **decreases**.

Hence, atomic radii decreases across the period.

▼ Down the Group

There is a general **increase** in atomic radius.

1. Number of filled principal quantum shells increases.
2. Distance between the nucleus and the valence electrons **increases**.
3. Electrostatic forces of attraction between the valence electrons and the nucleus **decreases** and the size of the electron cloud **increases**

Hence, atomic radii increases down the group.

▼ Group I elements (alkali metals)

▼ physical properties

1. soft and silvery
2. generally low densities and can float on water
3. low melting points
4. good conductors of heat and electricity

▼ chemical properties

1. they have 1 electron in their valence shell. their atoms react by losing one electron, forming an ion with a charge of $1+$
2. they are the **most reactive** of all the metals. Alkali metals are stored under oil to protect them from oxygen and water vapour in the air.
3. react violently with water — every alkali metal reacts with cold water, releasing **hydrogen** gas and producing a solution of metal hydroxide that is alkaline.

General equation: metal + water \rightarrow metal hydroxide + hydrogen gas

Example: 2 Na(s) + 2 H₂O(l) \rightarrow 2 NaOH(aq) + H₂(g)

observations:

1. effervescence is observed as hydrogen gas is produced
2. darting of the metal piece on the surface of the water.
3. beaker feels warm (heat is produced during the reaction)
4. colourless solution turns pink when phenolphthalein indicator is added to the solution
4. never found as free elements in nature
5. tend to react with non-metals forming soluble, white ionic compounds

▼ trends

▼ reactivity

reactivity of the elements **increases** down the group (metallic character increases).
The element shows a more violent reaction with water down the group.

Reason: Down the group,

1. **Number of filled principal quantum shells increases**
2. Distance between the nucleus and the valence electron **increases**.
3. Electrostatic forces of attraction between the nucleus and valence electrons **decreases**.
4. Thus, **less energy** is required to remove the valence electron from the atom.

In conclusion, first ionisation energy of the metals decreases. Thus reactivity of the metals increases.

▼ melting points

melting points of the elements **decreases** down the group.

▼ densities

the densities of the Group I Alkali metals are generally lower than other metals. As such, alkali metals such as lithium, sodium and potassium tend to float on water.

▼ Group VII elements (halogens)

▼ physical properties

1. typical non-metals with low melting and boiling points
2. they are all insulators of electricity.

▼ chemical properties

1. they have seven electrons in their valence shells
2. they all exist as **diatomic** covalently bonded molecules, with the formula X₂.
3. the atoms react by gaining one electron, forming an ion with a charge of 1-, X⁻
4. they form ionic compounds with non-metals (eg. hydrogen bromide (HBr), hydrogen chloride (HCl) and sulfur hexafluoride (SF₆)

▼ trends

▼ reactivity

Reactivity of the element **decreases** down the group

▼ melting and boiling point

Melting and boiling point **increases** down the group.

1. As the molecular mass of each molecule **increases**, the strength of the **intermolecular forces** of attraction (van der Waals' forces of attraction) **increases**.
2. **More energy** is needed to overcome the stronger intermolecular forces of attraction between the iodine molecules than those between bromine, chlorine and flourine molecules.

▼ colour

colour of the element gets darker down the group

▼ displacement reactions of halogens

Halogens are reactive non-metals. They react with most metals to form salts called halides. Fluoride, chloride, bromide and iodide ions are examples of halide ions

Halogens undergo displacement reactions with halide solutions. A displacement reaction is a reaction in which one element takes the place of another element in a compound

A **more reactive** halogen will displace a **less reactive** halogen from its compound

Note:

Flourine does not undergo displacement reactions

Flourine reacts violently with water to give hydrogen flouride gas (or a solution of hydroflouric acid) and a mixture of oxygen and ozone)

halogens in solution	KCl solution	KBr solution	KI solution
chlorine, Cl ₂ (aq) (light yellow)	no visible reaction	Solution turned from colourless to reddish-brown. Reasons: chlorine displaces bromine from potassium bromide solution	Solution turned from colourless to brown. Reason: chlorine displaces iodine from potassium iodine solution
bromine, Br ₂ (aq) (reddish-brown or orange)	no visible reaction	No visible reaction	Solution turned from colourless to brown. Reason: bromine displaces iodine from potassium iodine solution.
iodine, I ₂ (aq) (brown)	no visible reaction	No visible reaction	No visible reaction.

▼ noble gases (inert gases)

They are inert gases. Their inertness or lack of reactivity is due to their stable electronic configuration. They have complete valence shells. They have little tendency to share electrons to form a covalent bond or to lose or gain electrons to form an ionic bond. They are electronically very stable

▼ physical properties

1. they all are non-metallic elements
2. they are all colourless and odourless at room temperature and pressure, with low melting and boiling points

▼ chemical properties

1. They exist as monoatomic gases due to their electronic stability
2. they have high ionisation energies and low electronegativity values
3. helium is used for filling weather or advertisement balloons and argon is used to provide an inert atmosphere for processes such as welding stainless steel.

▼ transition metals

▼ physical properties

1. they have high melting and boiling points as well as high densities.
2. they are good electrical conductors

▼ chemical properties

1. they have variable oxidation states in their compounds. eg. the oxidation state of copper is +1 in Cu₂O and +2 in CuO
2. They formed coloured compounds. eg. aqueous iron (II) sulfate is pale green and aqueous iron (III) chloride is yellow. Compounds of transition metals are used in dyes, pigments, paints and stained glass to produce different colours
3. transition metals and their compounds are important catalysts for many reactions in the laboratory and industry. eg. nickel is used in the manufacture of margarine from vegetable. Iron is used in the Haber process for the manufacture of ammonia.

sample questions

▼ element X is in the same period as Z (group 1) and has two valence electrons.

Comparing the difference between X and Z, what would you observe when X is added to water? explain why.

Element X will react less violently with water compared to Z. Less effervescence is observed per unit time. This is because two valence electrons from X are removed to form X^{2+} ions in its reaction with H_2O while only one valence electron from Z is removed to form Z^+ ions in its reaction with H_2O . More energy is required to remove 2 valence electrons in X compared to one electron in Z. Hence, metal X will be less reactive than Z.

▼ **Name an element that can displace bromine. Give a reason for your choice.**

chlorine. A more reactive halogen will displace a less reactive halogen from its compound. Chlorine is more reactive than bromine as it is above bromine in the periodic table, and thus it can displace from the potassium bromide solution.

▼ **explain why the atomic radius of oxygen is smaller than that of beryllium**

Oxygen has 8 protons and beryllium has 4 protons, hence oxygen has a larger nuclear charge. As they are both in the same period, they have the same number of inner electrons. The electrostatic forces of attraction between valence electrons and the nucleus is greater for oxygen than beryllium and the size of the electron cloud/atomic radius is smaller for oxygen.

▼ **explain why the atomic radius of potassium is greater than that of sodium.**

Potassium atom has one more filled principal quantum shell than sodium atom. The distance between the nucleus and the valence electrons of potassium is greater than that of sodium. The EFOA between valence electron and nucleus is weaker for potassium and the size of the electron cloud/atomic radius is greater in the potassium atom.

▼ **which 2 elements are in the same group of the periodic table? explain.**

C and D. Both elements form oxides of similar chemical formulae. This shows that the elements formed ions that have a charge of 1^+ and they are both from group I of the periodic table as they have the same number of valence electrons.

Chemical Bonding and Structure

▼ **predicting bond type**

The **electronegativity** of an atom is a measure of its ability to attract the electrons in a covalent bond to itself. The smaller the atomic radius, the more rightly held the valence electrons are to the nucleus, the greater the electronegativity.

INSERT ELECTRONEGATIVE PIC

1. electronegativity increases across the period and decreases down the group
2. F is the most electronegative atom, followed by O and N

▼ **metallic bond**

the **strength of the metallic bond** is dependent on:

1. size of the cation — smaller size, stronger metallic bond
2. charge of the cation — larger charge, stronger metallic bond
3. no. of valence electrons — greater number of valence electrons delocalised in the lattice, stronger metallic bond.

▼ answering techniques

Name the structure, explain in terms of the chemical bonding, why metals have high melting point.

Because metals have a giant metallic structure where there are strong electrostatic forces of attraction between metal cations and ‘sea’ of delocalised electrons and a very large amount of energy is needed to overcome these strong electrostatic forces of attraction.

explain why metals can conduct electricity in solid and molten states.

The presence of mobile electrons that can act as **mobile charge carriers** allow metals to conduct electricity in solid and molten state

explain why metals are ductile and malleable

When a force is applied to the metal, the layer of positive ions slides over each other without disrupting the strong electrostatic forces of attraction between the cations and ‘sea’ of delocalised electrons

▼ ionic compounds

answering techniques

Name the structure, explain in terms of chemical bonding, why MgO has a high melting point.

Because ionic compounds have a **giant ionic lattice structure** where there are strong electrostatic forces of attraction between cations (mg ions) and anions (oxide ions) and a very large amount of energy is needed to overcome these strong electrostatic forces of attraction.

explain why ionic compounds can conduct electricity only in aqueous state or molten state but cannot conduct in solid state.

In solid state, the strong electrostatic forces of attraction between the cations and anions hold the ions in their fixed positions. There are no free ions to move and there are no mobile charge carriers to carry an electrical charge.

In molten, or aqueous state, the strong electrostatic forces of attraction are overcome and the ions are free to move and act as mobile charge carriers.

▼ simple molecular structure

Weak van der Waals' forces of attraction exist between molecules

Or if the molecule is eg. water, ammonia or hydrogen fluoride (polar molecules) — rem FON

There are **strong hydrogen bonds** between the molecules

Hydrogen bond definition: Hydrogen bond is an **electrostatic force of attraction** between an electron-deficient hydrogen atom covalently-bonded to a small **highly electronegative** atom (such as NOF) and the **lone pair electrons** of a nearby highly electronegative atom (N,O,F atom on another molecule)

*ADD HYDROGEN BOND IMAGE

intramolecular forces: strong covalent bonds

intermolecular forces (weak temporary dipole-induced dipole forces

1. van der Waals' forces
2. hydrogen bond

hydrogen bonds are stronger than van der Waals' between molecules but weaker than covalent, ionic, and metallic bonds.

lower density of ice compared to water

1. each H₂O molecule is hydrogen-bonded to four other H₂O molecules in a tetrahedral arrangement.
2. A very open structure with empty spaces between the water molecules accounts for ice being less dense than water at 0 degrees celsius.
3. When heat is applied to the molecules, the molecules gain kinetic energy to break the hydrogen bonds and cause the crystals to collapse. H₂O molecules in liquid state then pack more closely together in a smaller volume. Hence the water in liquid form has higher density at 0 degrees celsius. Water expands on freezing and contracts on melting.

▼ Explain, in terms of structure and bonding, why the discrete molecules have low boiling points/melting points eg. oxygen and methane molecules

Oxygen molecules have a simple molecular structure, with strong covalent bonds between the oxygen atoms.

Between the oxygen molecules, there are weak van der Waals' forces of attraction holding them together.

Small amount of energy is needed to overcome these weak van der Waals' forces of attraction.

▼ Explain, in terms of structure and bonding, why the discrete molecules have high boiling points (i.e. in comparison to oxygen molecules) e.g. water molecules

Water molecules have a simple molecular structure with strong covalent bonds between the oxygen atom and hydrogen atoms.

Between the water molecules, there are strong hydrogen bonds holding them together. A larger amount of energy is needed to overcome these strong hydrogen bonds.

▼ **giant molecular structure**

▼ **Explain, in terms of structure and bonding, why diamond has high melting points.**

Diamond has a giant molecular structure.

Each carbon atom is bonded tetrahedrally to four other carbon atoms by strong covalent bonds in a 3-dimensional network.

Large amount of energy needed to overcome numerous strong covalent bonds between the carbon atoms.

▼ **Explain, in terms of structure and bonding, why graphite has high melting points**

Graphite has a giant molecular structure.

Within each layer, each carbon atom forms strong covalent bonds with 3 other carbon atoms, leading to hexagonal rings of six carbon atoms that are joined together to form 2-dimensional flat layers.

Large amount of energy needed to overcome numerous strong covalent bonds between the carbon atoms.

▼ **Explain in terms of bonding and structure, why graphite can conduct electricity but diamond cannot.**

Diamond has a giant molecular structure.

Each carbon atom is bonded tetrahedrally to four other carbon atoms by strong covalent bonds in a 3-dimensional network.

There are no available valence electrons in the carbon atoms to act as mobile charge carriers.

On the other hand, for graphite also has a giant molecular structure.

Within each layer, each carbon atom forms strong covalent bonds with 3 other carbon atoms.

Therefore, each carbon atom has one free valence electron and these electrons are mobile along the same plane, allowing graphite to conduct electricity along the layers

▼ **Explain in terms of bonding and structure, why graphite can be used as a lubricant**

Graphite has a giant molecular structure.

Within each layer, each carbon atom forms strong covalent bonds with 3 other carbon atoms.

There are weak

van der Waals' forces of attraction between the layers and when a force is applied, these layers of carbon atoms can slide easily over each other.

▼ **The structure of this form of boron nitride resembles that of graphite, but boron nitride is an insulator while graphite is a conductor of electricity. Suggest why boron nitride does not conduct electricity.**

The valence electrons of each boron atom are involved in covalent bonding with 3 nitrogen atoms.

There are 2 valence electrons on every nitrogen atom that are not involved in covalent bonding. Nitrogen atoms are very electronegative and do not have mobile electrons.

Due to the absence of mobile charge carriers (mobile electrons), boron nitride does not conduct electricity.

▼ **alloy**

▼ **definition**

An alloy is a mixture of a metal with one or more elements

▼ **properties**

melting point: usually lower than pure metal

electrical conductivity: able to conduct

malleability and ductility: not as malleable and ductile as pure metal

In a metal, the regular and orderly arrangement allows the layers of atoms in a pure metal to slide over each other easily when a force is applied. Alloys have different physical properties from metals because the atoms of different elements are of a **different size** and disrupt the **regular** arrangement of atoms in the pure metal. In an alloy, the layer of atoms in alloys **cannot** slide over each other easily when a force is applied. Hence, an alloy is **stronger** than pure metal.

reasons for making alloys

1. to make metals stronger and harder
2. to improve the appearance of metals
3. to lower the melting points of metals
4. to make metals more resistant to corrosion

▼ **table**

	Metals	Ionic compounds	Simple molecules (covalent)	giant covalent substances
--	--------	-----------------	-----------------------------	---------------------------

	Metals	Ionic compounds	Simple molecules (covalent)	giant covalent substances
Structure	giant metallic lattice structure	giant ionic lattice structure	simple molecular structure	giant molecular structure
Bonding	<p>A metallic bond is the strong electrostatic forces of attraction between the ‘sea’ of delocalised valence electrons and positive metallic ions. Particles: positive metallic ions <u>and ‘sea’ of delocalised valence electrons</u></p>	<p>An ionic bond is the strong electrostatic forces of attraction between oppositely charged ions which formed as a result of the complete transfer of one or more electrons from one atom to another. Particles: cations and anions (name them)</p>	<p>A covalent bond is the strong electrostatic forces of attraction between the shared pair/pairs of electrons and the nuclei of both atoms. Usually, covalent bonds are formed between 2 non-metallic atoms whose difference in electronegativity is not too large Particles: <u>small, discrete molecules.</u> Answering technique: eg. Within each hydrogen chloride molecule, the hydrogen atom and chlorine atom are held by strong covalent bond.</p>	<p>Diamond: each carbon atom is bonded tetrahedrally to four other carbon atoms by strong covalent bonds in a 3-dimensional network forming a giant molecule. Graphite: Within each layer, each carbon atom forms strong covalent bonds with 3 other carbon atoms, leading to hexagonal rings of six carbon atoms that join together to form 2-dimensional flat layers . Weak van der Waals’ forces of attraction exist between the layers. Particles: carbon atoms</p>

	Metals	Ionic compounds	Simple molecules (covalent)	giant covalent substances
Electrical conductivity (good/poor)	<p>Metals are good electrical conductors in the solid and molten states. - metals have a giant metallic structure. It consists of positive ions surrounded by a 'sea' of delocalised electrons. The delocalised electrons acts as mobile charge carriers to conduct electricity under the influence of a potential difference. - conductivity increases from Na to Al as Al has the greatest number of delocalised electrons.</p>	<p>conduct electricity only in molten and aqueous states. but not in solid state. <u>In solid state</u>, strong electrostatic forces of attraction between oppositely-charge ions hold the ions in fixed position and are not able to move about freely. Hence, there are no mobile charge carriers to conduct electricity. In molten or aqueous states, the strong electrostatic forces of attraction between oppositely-charged ions are overcome. These ions are free to move around. Hence electricity can be conducted by these mobile ions.</p>	<p>Do not conduct electricity in <u>any state</u>. They are neutral molecules and do not contain mobile ions/delocalised electrons to act as mobile charge carriers.</p>	<p>Diamond does not conduct electricity - All the outer shell electron of the carbon atoms are used to form covalent bonds. Hence, there are no delocalised electrons to conduct electricity Graphite conducts electricity. - One valence electron per carbon atom is not used to form covalent bonds. These electrons are delocalised along the same plane and allow graphite to conduct electricity along the layers. (used as electrode in electrolysis)</p>

	Metals	Ionic compounds	Simple molecules (covalent)	giant covalent substances
melting point/boiling point (high/low)	- Metals have a giant metallic structure . - A large amount of energy is required to overcome the strong electrostatic forces of attraction between the cations and sea of 'delocalised' electrons .	High Large amount of energy is needed to overcome the strong electrostatic forces of attraction between oppositely-charged ions MgO has a higher melting point than NaCl — the greater the charge on the ions, the stronger the forces holding them together.	Low melting and boiling point Small amount of energy required to overcome the weak intermolecular forces between molecules .	high large amount of energy required to break the strong covalent bonds between atoms
physical state at r.t.p	solid (except for mercury)	solids	liquid/gas	solid
Solubility in water and organic solvents	Not soluble in any solvent	Soluble in water - water molecules (polar solvent) are polar molecules and are able to form ion-dipole bond with both positive and negative ions, which breaks up the lattice and causes the ions to separate. Insoluble in organic solvent. - Non polar. They do not interact with ions in the ionic compound and hence are not able to break up the ionic lattice.	Insoluble in water (polar solvent) exceptions: However, water can interact with and dissolve some smaller molecules, eg. sugar, ethanol and acids. This is due to the ability of these molecules to form hydrogen bonds with each other. Soluble in organic solvent (non-polar)	INSOLUBLE

	Metals	Ionic compounds	Simple molecules (covalent)	giant covalent substances
other properties	<p>Ductile and malleable Since the valence electrons in the lattice do not belong to any particular metal atom, if sufficient force is applied, one layer of positive ions can slide over another without disrupting the metallic bonding.</p>	<p>Hard and brittle When a force is exerted on the ionic compound, ions with similar charge from adjacent layers face each other.</p> <p>Repulsive forces between the similarly charged ions splits the crystals</p>		<p>Diamond: very hard - it is difficult to break the numerous strong covalent bonds between the atoms present throughout the giant molecular structure. [used as drill tips for drilling equipment and in glass cutting]</p> <p>graphite: soft The layers of carbon atoms are held loosely by van der Waals' forces of attraction. These layers of carbon atoms can slide over each other when a force is applied. [used as a solid lubricant to reduce friction in engines, and also used as pencil lead]</p>

sample questions

▼ structure and bonding

type of structure	description of type of bonding
giant molecular structure	strong and extensive covalent bonds between atoms
giant ionic lattice structure	strong electrostatic forces of attraction between cation and anion/oppositely charged ions
simple molecular structure	strong covalent bonds between atoms; weak van

	der Waals' forces between molecules
giant metallic structure	strong electrostatic forces of attraction between positive metallic cations and 'sea' of delocalised valence electrons