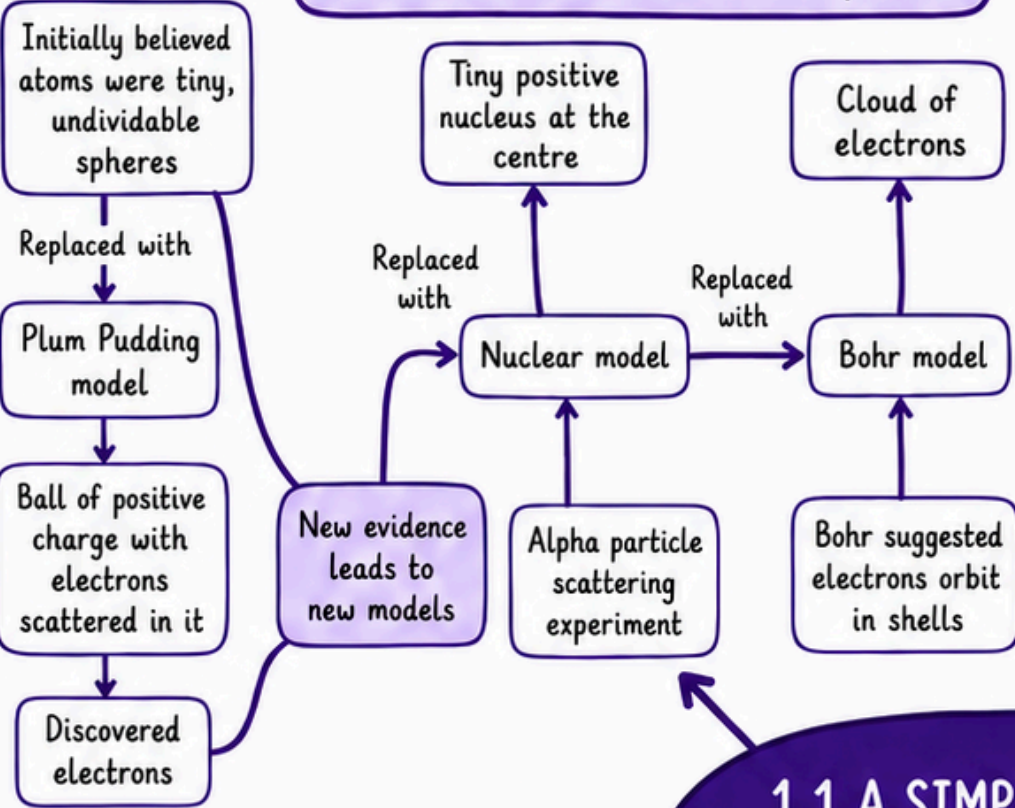
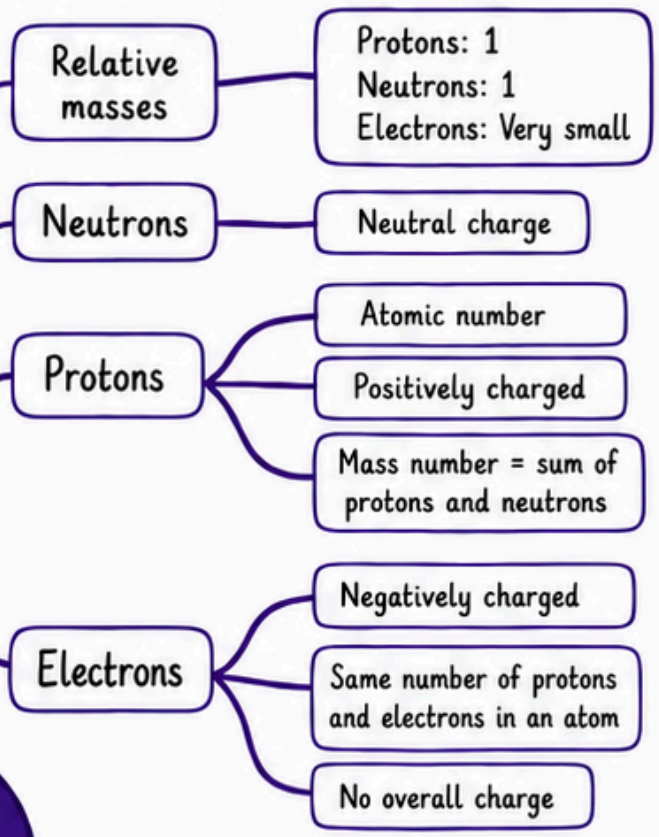




## 1. Atomic model development

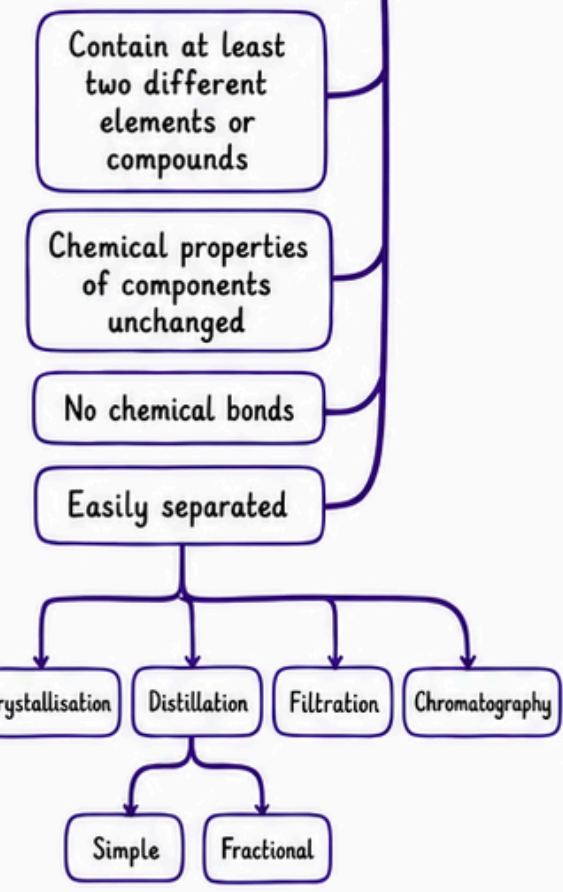


## 2. Subatomic particles

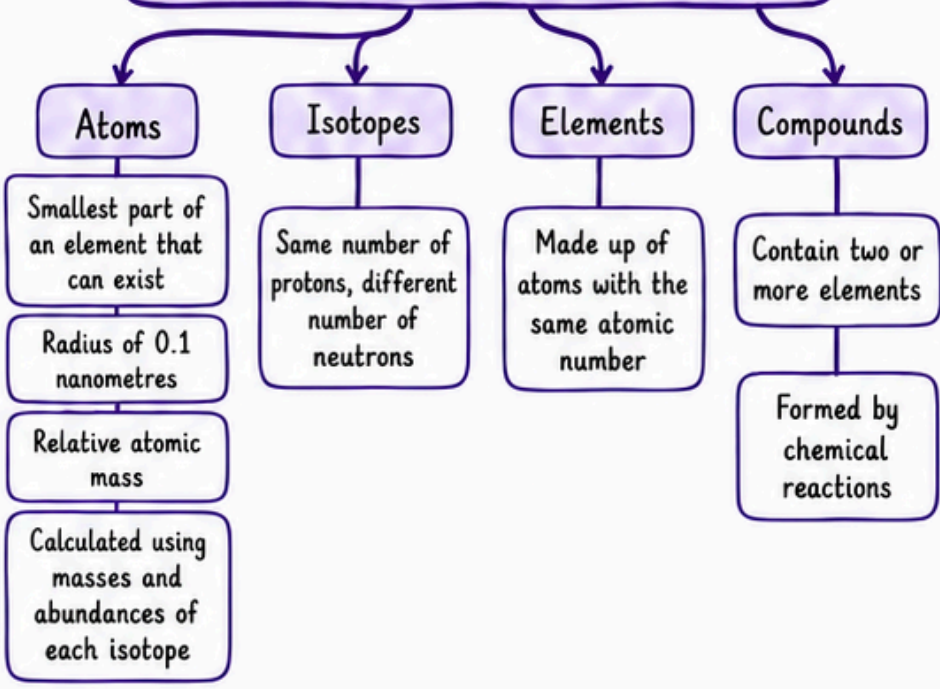


# 1.1 A SIMPLE MODEL OF THE ATOM, SYMBOLS, RELATIVE ATOMIC MASS, ELECTRONIC CHARGE AND ISOTOPES

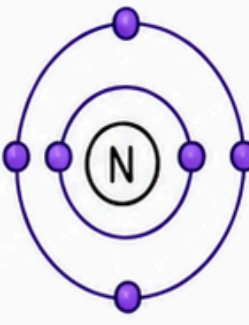
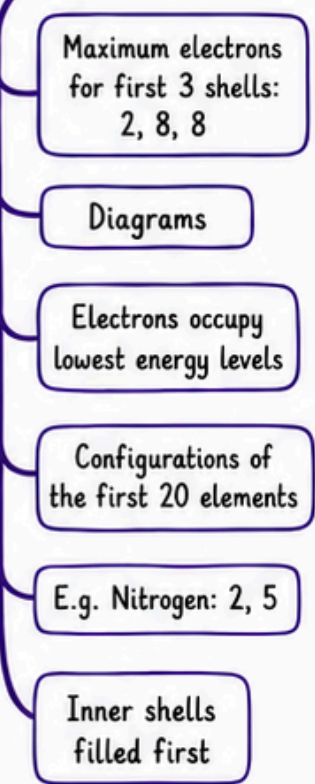
## 5. Mixtures

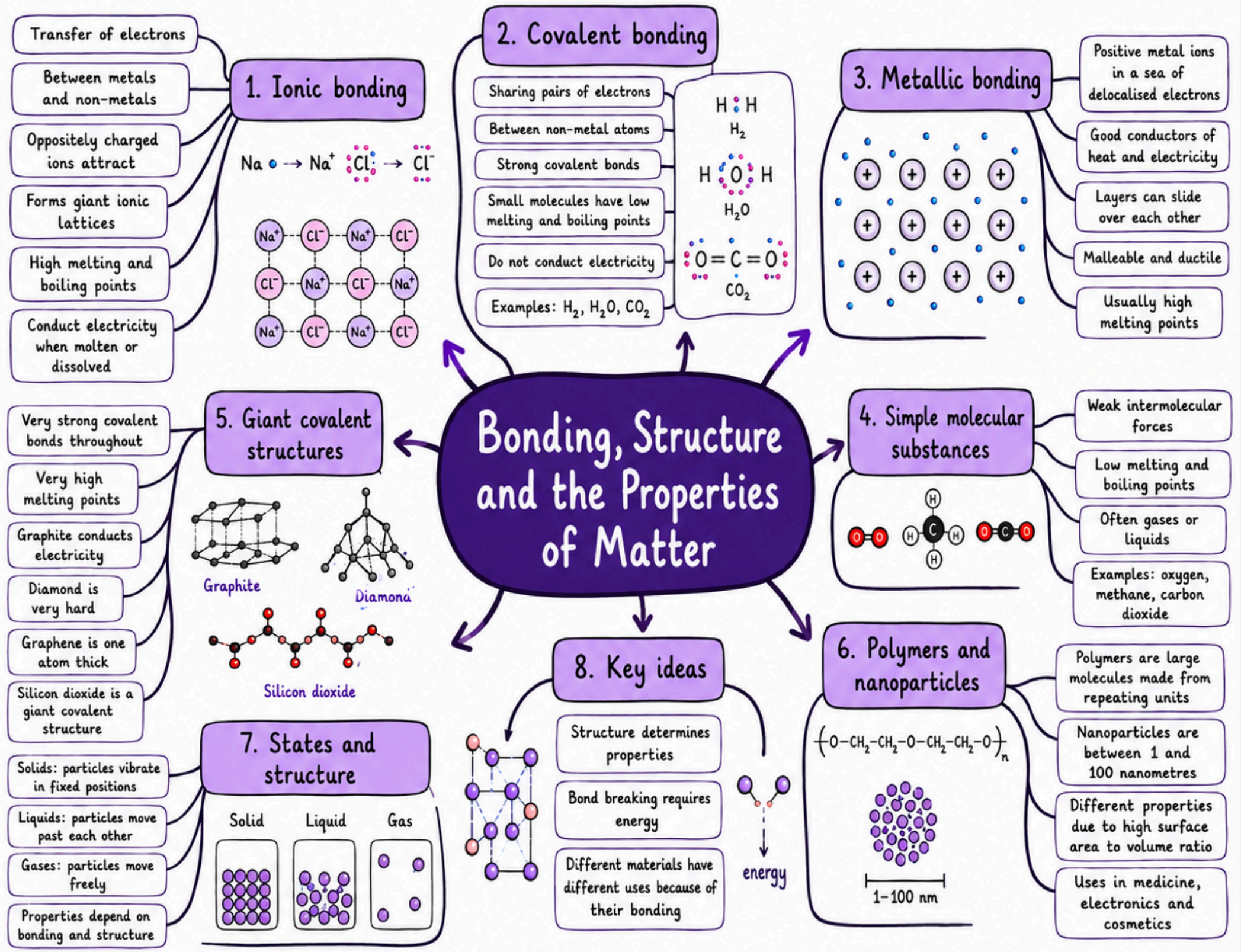


## 4. Atoms, Elements and Compounds



## 3. Electronic structure





**Key**

- = electron from atom A
- = electron from atom B
- = other atom

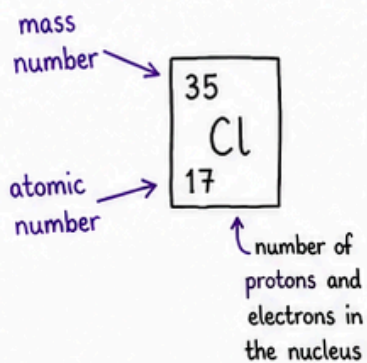
Electron sharing (covalent)	Ionic lattice (example: NaCl)	Metallic structure	Particle diagrams						
$\text{H} \cdot \text{Cl} \cdot$			<table style="width: 100%; text-align: center;"> <tr> <th>Solid</th> <th>Liquid</th> <th>Gas</th> </tr> <tr> <td></td> <td></td> <td></td> </tr> </table>	Solid	Liquid	Gas			
Solid	Liquid	Gas							

**Remember:**  
Strong bonds = high melting points  
Weak forces = low melting points





### 3.1 - Relative Formula Mass



**Relative atomic mass (Ar)**  
average mass of all the isotopes of an element.

**Isotopes** = same amount of protons, different amount of neutrons.

#### Relative formula mass (Mr)

add together the relative atomic masses (Ar) of all the atoms in the compound or molecular formula.

E.g. magnesium chloride



$$24 + (35.5 \times 2)$$

$$= 24 + 71$$

$$= 95$$

**Mr = 95**

#### Percentage mass

the percentage of mass of a particular element in a compound.

E.g. % mass of sulfur in  $H_2SO_4$

$$\frac{(32 \times 1)}{98} = 0.327$$

$$\times 100$$

$$= \mathbf{32.7\%}$$

### 3.2 - Moles and mass

#### Mole

unit used to measure amount of chemical we have.

1 mole of any substance contains  $6.02 \times 10^{23}$  particles (atoms, molecules, electrons etc.).

↓  
Avogadro's constant

The mass of  $6.02 \times 10^{23}$  particles of any substance equals its relative formula mass in grams.

E.g.  $O_2$ , 1 mole = 32 g  
Mr = 32

#### Moles formula

$$\text{number of moles in a sample} = \frac{\text{mass of that element compound (g)}}{\text{Mr}}$$

### 3.3 - Calculating Mass in reactions (Example)

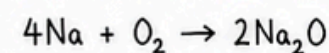
Calculate the mass of sodium oxide ( $Na_2O$ ) that can be made by completely burning 5.0 g of sodium (Na) in oxygen. Give answer to 3 sig. fig.

Known substance  
= sodium (Na)

Unknown substance  
= sodium oxide ( $Na_2O$ )

#### The main 4 steps

- ① Write out balanced equation
- ② Find the moles of known substance (moles =  $\frac{\text{mass}}{\text{Mr}}$ )
- ③ Find the moles of unknown substance using molar ratio from balanced equation
- ④ Find the mass of the unknown substance



• moles of Na =  $\frac{5}{23} = 0.217$  (step 2)

• molar ratio  $4Na : 2Na_2O$   
simplifies to 2 : 1 (step 3)

• moles of  $Na_2O = 0.217 \div 2 = 0.1085$  moles

• Mr of  $Na_2O = (23 \times 2) + 16 = 62$  (step 4)

• mass of  $Na_2O = 0.1085 \times 62 = 6.7279$

**To 3 sig. fig., mass of  $Na_2O = 6.73$  g**

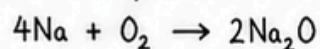
### 3.5 - Limiting reactants

**Limiting reactant** is the reactant that limits the amount / mass of product produced.

**Other reactant** is in excess.

Example: calculate mass of sodium oxide produced when 115 g of sodium is burned in air.

① Balanced equation (combustion / oxygen in air is in excess)



② Sodium in moles

$$\frac{115}{23} = 5 \text{ moles Na}$$

③ Ratio Na :  $Na_2O = 2 : 1$

5 moles Na gives 2.5 moles  $Na_2O$

④ Mass of  $Na_2O$

$$\text{mass} = \text{moles} \times \text{Mr}$$

## Quantitative chemistry

### 3.4 Conservation of Mass

No atoms are created or destroyed during a chemical reaction, only the bonds change.

$$\text{Relative formula mass of reactants} = \text{Relative formula mass of products}$$

$$\text{Actual mass (g) of reactants} = \text{Actual mass (g) of products}$$

Exception: when reactions involve a gas, mass may seem to change if gas enters or escapes.

↑ Increase in mass if one reactant is a gas.

↓ Decrease in mass if one of the products is a gas.

