Elements: Only one type of atom
Compound: More than one type of atom, chemically bonded
Mixtures: more than one type of atom, not chemically bonded

Separation techniques: **filtration, crystallisation, chromatography and distillation** (solution is heated and the solvent vapour evaporates. Is condenses and can be collected)

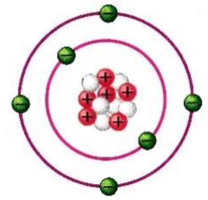
Atom theory
"tiny spheres" → electrons discovered → plum pudding model → alpha particle scattering experiments → nuclear model (Bohr)→ proton discovered → Neutron discovered (Chadwick)

Alpha Scattering experiments
Positive particles shot through gold atoms. Most of the particles went straight through so atom not solid and mostly empty space (plum pudding wrong). Some positive particles deflected by the small positive nuclei in the centre (nuclear model correct)

The **periodic table** was initially arranged by **atomic weight** meaning atoms were in inappropriate groups (columns)

Elements with properties predicted by **Mendeleev** were discovered and filled the gaps he had originally put into his table.
Metals are bottom left and **non-metals** are top right

Sub- atomic particle	Charge	Mass
Proton	+1	1
Electron	-1	0
Neutron	0	1



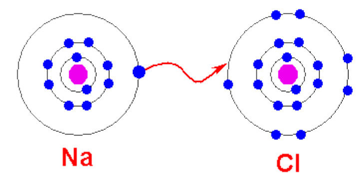
C1

Atom Radius = 10.1nm or 1 x 10^{-10m}
Nucleus Radius =1/10,000 or an atom radius or 1 x 10^{-14m}

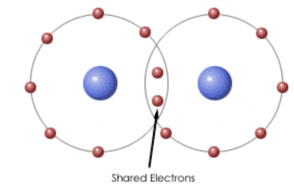
Strong bonds = High melting (liquid at room temp) and boiling points (gas at room temp) because it takes a **lot of energy** to break the strong bonds and for substance to change state

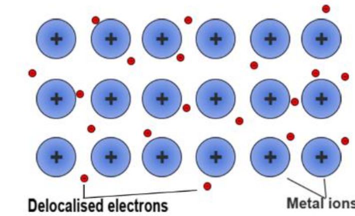
Mass number = protons + neutrons
Atomic number = protons (same as electrons)
Isotopes = different versions of an element with the same protons but different neutrons
Positive ions lose electrons, **negative ions** gain electrons

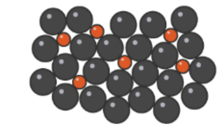
Ionic bond – metal + non-metal
Gain and lose e-



Covalent bond – 2 non-metals
Share e-

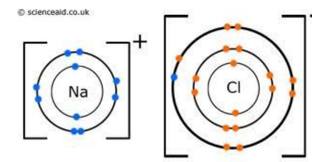






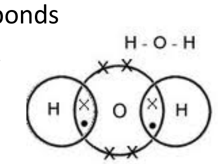
Ionic compounds: metal + non-metal e.g. Group 1 alkali metal and Group 7 Halogen
Positive ions lose electrons, negative ions gain electrons – ions have full outer shell (like Group 0)

Ionic compound is a **regular giant lattice** of ions
- **Strong electrostatic forces** of attraction between oppositely charged ions.
-**High mp/bp** because of the large amounts of energy needed to break the **many strong** ionic bonds.



-**Conduct electricity** when molten or dissolved because the ions are free to move and carry the current

Covalent compounds: 2 non-metals **share** pairs of electrons to form strong covalent bonds
Small covalent molecules have just a few atoms e.g H_2, Cl_2, O_2, HCl, H_2O, NH_3 and CH_4
- **Low mp/bp** due to weak **intermolecular forces**
- Do not conduct electricty as do not have any charged particles

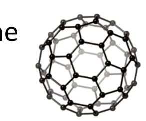


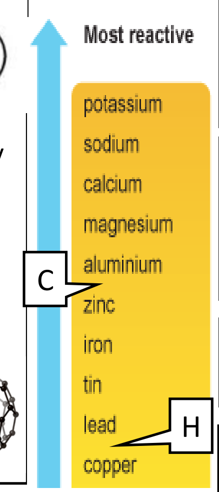
Giant covalent molecules: Giant lattices, **many strong** covalent bonds so **high mp/bp**

Diamond: each carbon atom forms **4 bonds** with other carbons = very hard, does not conduct electricity (Silicon dioxide has similar structure but made of silicon and oxygen)

Graphite: each carbon atom forms **3 bonds** with other carbons, forming **layers of hexagonal rings**
-The **layers slide** over each other because there are no covalent bonds between the layers and so graphite is **soft and slippery** (weak intermolecular forces).
- One electron **from outer shell** of each carbon atom is **delocalised =** conduct heat and electricity.

Fullerenes and carbon nanotubes: Carbon molecules with hollow shapes .The first fullerene discovered was **Buckminsterfullerene (C60)**. High SA:V ratio. Good for nanotechnology and electronics. **Graphene** is a single layer of graphite – electronics and composites





Metal -Good conductor of heat and electricity and malleable

Metallic bonding - The electrons in the outer shell of metal atoms are **delocalised** and so are free to move through the whole structure (carry the charge). There is a strong electrostatic force and so metals have high melting and boiling points

Alloys -Mixture of metal and another element e.g. steel
Stronger than pure metal because the different size atoms distort the layers so they can no longer slide

Metal + Oxygen → metal oxide
Metal + acid → metal salt + hydrogen
Metal oxide + acid → metal salt + water
Metal carbonate + acid → metal salt + water + carbon dioxide

Metals less reactive than carbon can be extracted from their oxide **ores** by **reduction with carbon**. Oxidation involves the gain of oxygen. Reduction involves the loss of oxygen
e.g. iron oxide + carbon → iron + carbon dioxide

Hydrochloric (HCl) =chloride, Nitric (HNO_3) = Nitrate, Sulfuric (H_2SO_4) = Sulfate

To make a **soluble metal salt** the solid is added to the acid until no more reacts and the excess solid is filtered off to produce a solution of the salt.

Salt solutions can be crystallised to produce solid salts

Group	Properties	Equations
0 **Noble gases**	**Unreactive** due to full outer shells **Boiling points increase** has the atoms get bigger – larger atoms will have stronger intermolecular forces	Unreactive
1 **Alkali metals**	1 electron in outer shell **Reactivity increases** going down the group – larger atoms, the outer electron further away from nucleus so lost easily in a reaction	Alkali metal + water → metal hydroxide + hydrogen Alkali metal + oxygen → metal oxide Alkali metal + chlorine → metal chloride
7 Halogens	7 electrons in outer shell – non metals, pairs of atoms Melting and boiling points increase down the group **Reactivity decreases** going down the group – larger atoms, the outer electrons further away from nucleus so hard to gain electrons (less pull)	Metal + halogen → metal halide (e.g. tin chloride, tin bromide, tin iodide) Non-metal + halogen→ non-metal halide

Displacement: A more reactive halogen can displace a less reactive halogen e.g. chlorine + potassium iodide → potassium chloride + iodine

PH scale: 1 = strong acid, 14 = strong alkali, 7 = Neutral

H+(aq), make solutions acidic
OH–(aq), make solutions alkaline

acid + alkali = neutralisation
$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$ [l liquid, g gas, s solid, aq aqueas]

Strong acids are **completely ionised** in solution (hydrochloric, sulfuric, nitric)

Weak acids are **partially ionised** in solution (citric, ethanoic, carbonic)

At the same concentration there are more H+ in a strong acid. As the pH decreases by one unit, the H+ concentration of the solution increases by a factor of 10

Electrolysis
Splitting a dissolved or molten ionic compound (electrolyte) using electricity (expensive, uses energy)

Negative ions move to the positive electrode and lose e- to become discharged atoms (**Oxidation** is losing OIL)

Positive ions move to the negative electrode and gain e- to become discharged atoms (**Reduction** is gaining RIG)

We use electrolysis to extract metals that are more reactive than carbon.
Aluminium extraction from aluminium oxide(cryolite lowers the mp to make process cheaper)
Aluminium forms at the negative electrode and **oxygen** at the positive electrode. The positive electrode is made of carbon, which reacts with the oxygen to produce **carbon dioxide** – it must be replaced regularly

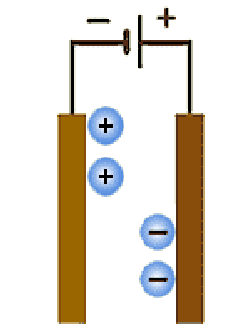
If the ionic compound is in a aqueous solution there will also be H+ and OH-

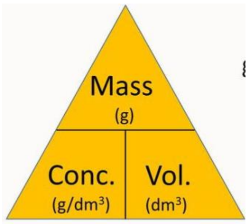
The product at each electrode follow these **rules**:
Positive electrode – If a halogen is present that is produced, if any other negative ion then oxygen produced
Negative electrode – If it is a very reactive metal then hydrogen is produced e.g. Sodium. If it is a metal with low reactivity then that metal is produced e.g. Copper

Electrolysis of sodium chloride solution (Brine) produces **hydrogen** (used for fuel) and **chlorine** (used for bleach and plastic) **Sodium hydroxide** solution is left behind (used for soap).

Half equations: $2H^+ + 2e^- \rightarrow H_2$
$2Cl \rightarrow Cl_2 + 2e^-$ or $2Cl^- - 2e^- \rightarrow Cl_2$

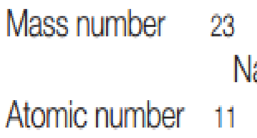
Step 1: Check number of ions and atoms the same
Step 2: check the number of e- correct





Bias:
Organisation or companies could benefit from presenting data in a certain manner. It is better to use an "independent" company or scientist to avoid bias (they have nothing to gain)
Peer reviews make scientific results more trustworthy

Calculations
Ar: relative atomic mass of an element compared to the mass of atoms with the ^{12}C isotope
It is an **average value** for the isotopes of the element (look up mass number)



Mr: relative formula mass of a compound is the sum of the relative atomic masses
e.g $CaCO_3$ 40 + 12 + 16 + 16 + 16 = 100 **1 mole = Mr in grams**

The number of atoms, molecules or ions in a mole of a given substance is the Avogadro constant.
The value of the **Avogadro constant is 6.02 x 10^{23} per mole**

If you wanted to calculate the **number molecules** in 4 mol of water you would calculate 4 x 6.02 x 10^{23}
If you wanted to calculate the **average mass of an atom/molecule** you would calculate Mr/Avogadro constant
Average mass of a $CaCO_3$ molecule = 23g/6.02 x 10^{23}

The mass of reactants always is the same as mass of products (**law of conservation of mass**)
If it seems that this has not happened it could be some product was escaped as a gas or it could be a reversible reaction

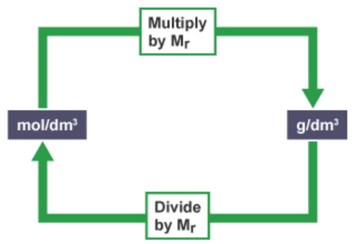
Reacting Mass calculation: Calculate the mass of MgO made if you have 10g of Mg

Mg + O → MgO
24 → 40
1g of Mg → 40/24 = 1.67g
10g of Mg → 16.7g of MgO

Step 1: Put in the Mr of the substances mentioned in Q
Step 2: simplify to 1g
Step 3: multiply up for the mass given in Q

In a reaction we often add an excess of one of the reactants so that all of the other reactant used up – the reactant used up is called the **"limiting reactant"** as it limits to amount of product

Concentration of solution = mass of solute in g/ volume in dm³ (g/dm³)
e.g. If you have 10g of solute in 100dm³ = 0.1 g/dm³
You can also express in mol/dm3

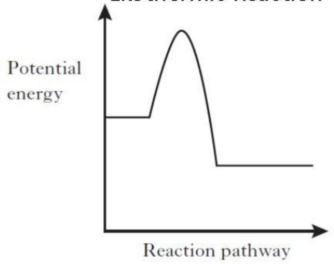


RP1: Making salts: To make a **soluble metal salt** the solid is added to the acid until no more reacts and the **excess** solid is filtered off to produce a solution of the salt. Salt solutions can be crystallised to produce solid salts e.g. $CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$
RP2: Electrolysis: State the ions in the aqueous solution or use the "rules" to predict the products at each electrode
RP3: Temperature Change: Read the Q carefully to see which variable is being changed e.g. IV: Mass of one of the reactants, DV: Temperature change, CV: Concentrations of reactants, Volume of reactants, Time, temperature of the room, surface area of the beaker

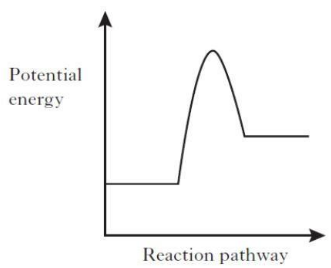
Energy changes
Exothermic → energy released to surroundings in reaction (combustion, neutralisation, oxidation)
Endothermic → energy taken in from surroundings (photosynthesis, thermal decomposition, citric acid + sodium hydrogencarbonate)

Activation energy: Energy required for particles to collide with enough energy to react – shown by "hump" on reaction profiles

Exothermic Reaction



Endothermic Reaction



Energy absorbed to break bonds, energy released when bonds form

Exothermic – **energy released** to form> **energy absorbed** to break = energy released overall
Endothermic - **energy released** to form< **energy absorbed** to break = energy absorbed overall

Designing experiments
Valid experiments answer the original question – e.g. if you want to test if mobile phones give you cancer, there is no reason to test the effect of pollution

- **Independent variable: Changed**
- **Dependent variable: Measured**
- **Control variables: Kept the same**-Experiments must be **fair**

If you want to do an experiment at 20cm, 30c,m , 40cm the **range** is 20-40cm and **interval** is 10cm
Equipment (apparatus) should have the correct **resolution** (smallest change that can be detected) e.g. It I need to measure 12ml of liquid I can not use a beaker that has intervals of 50ml – it is not accurate enough
The **uncertainty** of a measurement is half the resolution – e.g. a measuring cylinder with 10ml resolution has an uncertainty of + or – 5ml

Bond energy calculations
e.g. H-H = 436, Cl-Cl = 243, H-Cl = 432 (KJ/mole)
H–H + Cl–Cl → 2 × (H–Cl)
432 = 243 = 679 → 2 x 432 = 864
Energy change = in – out = 679 – 864 = -185 KJ/mole

Negative energy change = exothermic
Positive energy change = endothermic

Random errors could include not starting and stopping the stopwatch at exactly correct time – remove anomaly and repeat
Systematic errors is when the same error is made every time – whole experiment needs to be repeated
Zero Error – Scales have a reading even if nothing is on them. You should restart scales or data logger or take away the error for each result

Conclusions
DESCRIBE – give the pattern and include data
EXPLAIN – give the science to explain why you see this pattern
EVALUATE – give evidence for and against