Chapter 1 - Matter and Structure of Atoms

> Key point of the chapter:
>
> - Classification of Matter(states, particles)
> - *Change of State
> - Structure of Atoms(History, Electronic *Structure, Isotopes)
> - *Chemical Formulae(Nomenclatures, Oxidation Numbers)
> - Physical and Chemical Properties
> - Separation of Mixtures

Classification of Matter

States of Matter

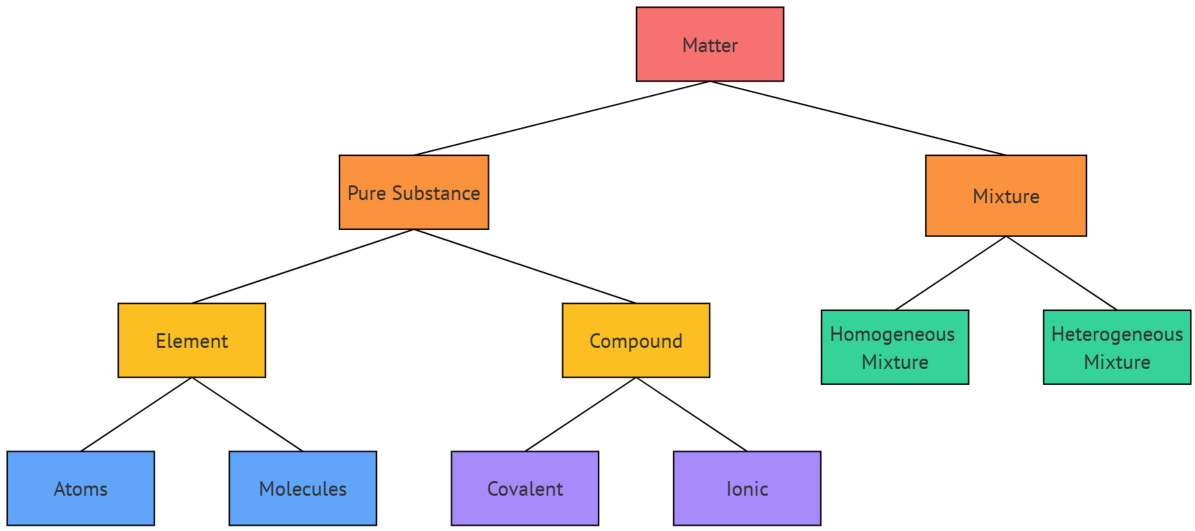
- Matters have **volume** & **mass**
- Matter can exist in three different **physical states**: **solid**, **liquid**, **gas**
- The states of matter can be changed by changing the **temperature** and/or **pressure**
  - Changes in pressure of temperature have greater impact on **gas**

Physical State	solid	liquid	gas
Volume	fixed volume	fixed volume	no fixed volume
Density	high	moderate to high	low
Shape	fixed shape	no fixed shape	no fixed shape
Fluidity	does not flow	generally flows	flows easily
Movement	vibrate in a fixed position	free to move	move freely
Particles	closely packed	slightly further apart	far apart from one another
Attraction Force & Kinetic Energy	$E_k < F$	$E_k \approx F$	$E_k > F$

- **Attraction forces (F)** between the particles won't change, but the kinetic energy (E_k) of particles will be affected by heat.

Particles

- There're three types of particles that made up matters: **atoms**, **molecules**, **ions**

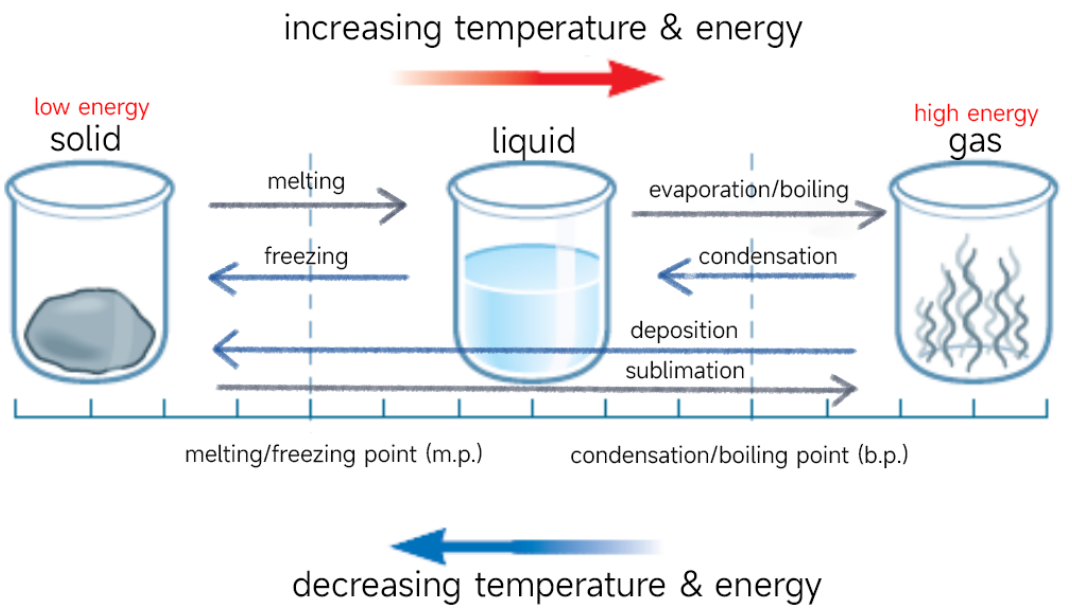


Difference between Compounds and Mixtures

Compounds	Mixtures
A compound is a single substance.	A mixture contains two or more substances.
The composition is always the same.	The composition can be varied.
The formation involves a chemical reaction.	No chemical change takes place when made.
The properties are very different from the elements present in the compound.	The properties of the substances making the mixture are still present.
Can only be broken down by chemical reactions.	The substances present can be separated by physical methods.

The Change of State of Matter

- All states of matter show an increase in volume (**expansion**) when the **temperature** is **increased** and a decrease in volume (**contraction**) when the **temperature** is **lowered**
- Large **increases/decreases** in **temperature** and **pressure** can cause a substance to **change** its **physical state**



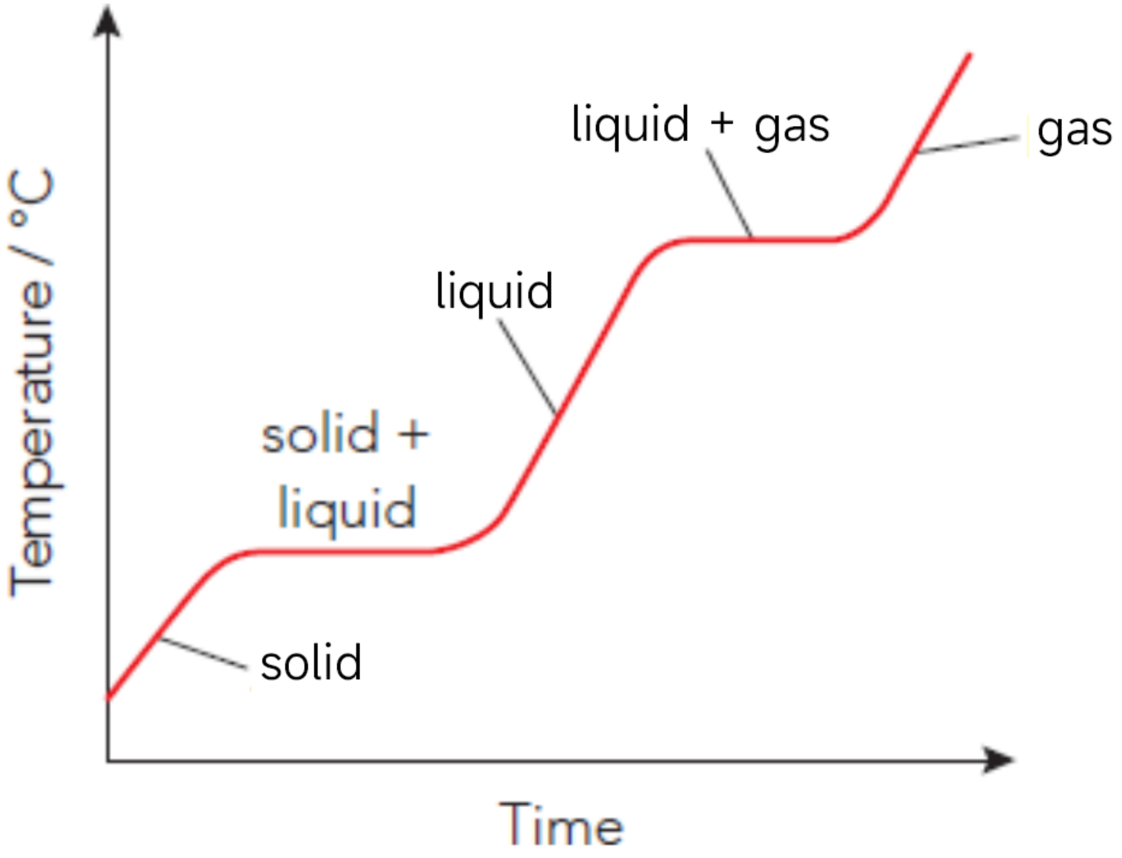
Boiling & Evaporation

- **Evaporation** take place over a range of temperatures, only take place from the surface of the liquid
    - the rate of evaporation may be affected by the **temperature** and **surface area**
    - *volatile*: term that describes a liquid that evaporates easily
        - usually a liquid with low boiling point because there're weal intermolecular forces between the molecules in the liquid
- **Boiling** takes place at a **specific temperature**

Sublimation

- Example of substances that sublimes:
    - iodine, dry ice (solid CO_2), mothballs, ammonium chloride (NH_4Cl)
- Ammonium chloride doesn't REALLY sublime, at a certain temperature, it decomposes to ammonia and hydrogen chloride (both gases), which appears to be sublimation but is not.
- Iodine can exist in liquid state. The iodine seems to miss out the liquid stage if heated strongly because the temperature rises very quickly, the m.p. and b.p. are close together, the liquid stage is not seen as it boils quickly
    - To make iodine melt to form a liquid at atmospheric pressure:
        - heat the solid slowly
        - use some tools (electrical heater, oil bath) to control the rise of temperature carefully

Heating and Cooling Curves



- The **temperature stays constant** during the state changes.
- During heating, all the energy absorbed at m.p. and b.p. goes into weakening the attraction forces between particles without temperature rise.
- During colling, all the heat energy released at condensation point and freezing point is given out into increasing of the intermolecular forces without temperature falls.
- *exothermic changes*: freezing, condensing
- *endothermic changes*: melting, boiling, evaporation

Effect of impurities

- **A pure substance melts and boils at definite temperature**.
- The presence of any **impurity** will **lower the freezing/melting point** of the solid
    - It makes it harder for the particles to arranged closely in fixed positions which is necessary if a substance is changed into a solid
    - It disrupts the formation of the lattice of the solid, less energy is required to break the lattice, this leads to a lower melting point
    - *lattice*: a regular 3D arrangement of particles in a crystalline solid
- The presence of any **impurity** will **increase the condensing/boiling point** of the liquid
    - It makes it harder for the particles to gain high kinetic energy which is necessary if a substance is changed into a gas/vapour(mixture of two states)
    - the molecules need more energy(heat) than normal in order to escape from the liquid and become vapour.
- Thus, **purity of substance** is checked by **determining the m.p. and b.p.**
- These values can also be used to **identify an unknown substance**.

Structure of Atoms

The Historical Development of Atomic Model

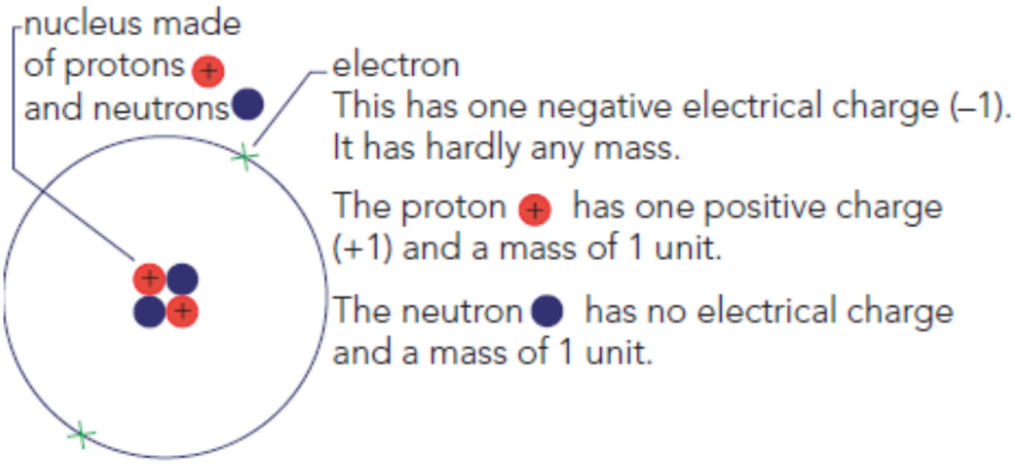
1. John Dalton
    - atomic model: all matter was made up of atoms
2. J.J. Thomson
    - plum pudding model: discovered electron(negatively charged particles)
3. Ernest Rutherford
    - Rutherford model: discovered protons(positively charged particles) and nucleus
4. Niels Bohr
    - Bohr model / planetary model: electrons move in specific energy levels or "shells" around the nucleus of an atom
5. James Chadwick
    - discovered the neutron(electrically neutral / no electrical charge)
    - completing the picture of the three main *subatomic particles* that make up atoms

Characteristics of Protons, Neutrons and Electrons

Subatomic Particle	proton(p)	neutron(n)	electron(e)
Relative Mass	1	1	1 / 1840 (negligible)
Relative Charge	+1	0	-1
Location in Atom	in nucleus	in nucleus	surrounding/outside nucleus

- **Atoms are electrically neutral**, so every atoms have an **equal number of protons and electrons** to cancel out the charges
- The **number of neutrons** required to hold the nucleus together **increases** as the **atomic size (atomic number) increases**
    - **Hydrogen** atom is the only atom that has **no neutron**
    - Neutrons are important in making the nucleus of an atom stable
        - Protons are positively charged and would therefore repel each other.
        - The presence of the neutrons counteracts this repulsion and means that the nucleus can hold together.

Example: Helium



A helium atom has these charged particles in it:

2 protons	charge +2	these charges cancel out
2 electrons	charge –2	

We say the charges balance.
The atom has no overall electrical charge.

A helium atom has:

2 protons	mass 2 units
2 neutrons	mass 2 units
2 electrons	with hardly any mass

So a helium atom has a total mass of:
2 + 2 = 4 units

Atomic Number and Mass Number

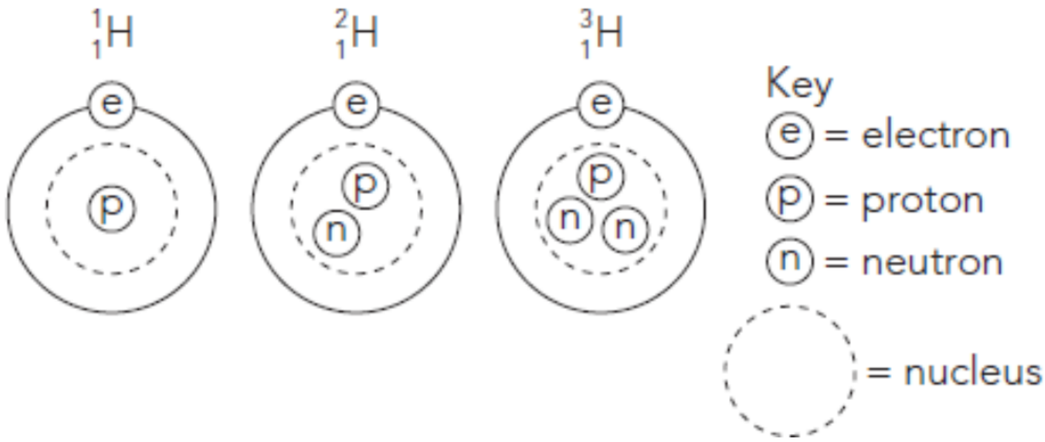
- We can represent a particular atom of an element by combining the chemical symbol of the element with the nucleon(A) and atomic(Z) numbers of the atom.

$$^{A}_{Z}\mathrm{X}$$

- Example: representing the structure of iodine atoms, using the above format
  - $$^{127}_{53}\mathrm{I}$$ $$Iodine - 127$$
- Atomic number / proton number (Z) = number of protons in the nucleus of an atom
- Nucleon number / mass number (A) = number of protons + number of neutrons
- number of electrons = number of protons = atomic number
- number of neutrons = mass number - atomic number (A-Z)

Isotopes

- *Isotopes* are atoms of the same element that have the **same proton number** but a **different nucleon number**. Many Isotopes are **radioactive** because their nuclei are unstable (radioisotopes).
- Many elements have naturally occurring isotope.
- Hydrogen, the simplest element, has two naturally occurring isotopes: hydrogen and deuterium. A third Isotope, tritium, can be made artificially.



- The isotopes of an element have the **same chemical properties**.
  - Isotopes have the **same electronic configuration**. It determine the way it forms bonds and react with other atoms.
- The isotopes **differ in some of their physical properties** (such as density, rate of diffusion)
  - Examples
    - difference in density between ordinary ice and heavy-water ice(deuterium oxide, D_2O)
    - radioactivity of tritium and carbon-14 (radioisotopes)

Relative Atomic Mass (A_r)

- A single atom cannot be weighed on a balance.
- Masses of all atoms are compared to the mass of a **carbon-12 atom**.
- One atom of carbon-12 is given the mass of 12 precisely.
- $$\text{1 atomic mass unit (a.m.u)} = \frac{1}{12} \times \text{mass of one atom of carbon-12}$$
- Examples of value obtained for other element
  -

Element	Atomic Symbol	Relative Atomic Mass (A_r)
carbon	C	12
hydrogen	H	1
copper	Cu	64

- **Most elements exist naturally as a mixture of isotopes**, therefore, the value we use for the **atomic mass** of an element is an **average mass**.

$$A_r(X) = (\text{isotope}_1 \text{ abundace} \times \text{mass of isotope}_1) + (\text{isotope}_2 \text{ abundace} \times \text{mass of isotope}_2)$$

- Examples
  - Iridium has two isotopes. These isotopes are iridium-191 and iridium-193. A natural sample of iridium consists of 37.3% of iridium-191.
  - $$\begin{aligned} A_r(Ir) &= (191 \times 37.3\%) + (193 \times (100 - 37.3)\%) \\ &= (191 \times 37.3\%) + (193 \times 62.7\%) \\ &= 192.254 \end{aligned}$$

Electronic Configuration of Elements

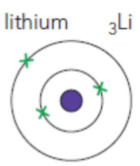
Bohr's atomic theory

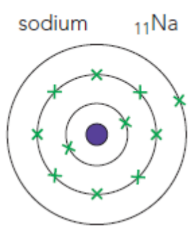
- electrons are in **orbit** around the central nucleus of an atom
- the electron orbits are called **electron shells / energy levels** and have different energy
- **shells that are further from the nucleus have higher energies**
- the shells are filled starting with the one with lowest energy (closest to the nucleus)
- the first shell can hold only two electrons
- the outermost shells can hold eight electrons to give a stable arrangement of electrons
- If the number of electrons of an atom is more than 20, the third shell will hold 18 electrons. If the numbers of electrons is 20 or less, the third shell will hold 8 electrons
- **electron arrangement / configuration**: the way in which the electrons are distributed in the shells of an atom

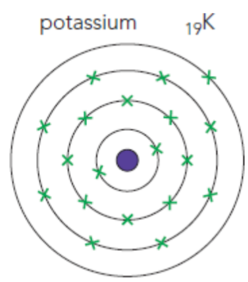
Example of Electron Configurations of The First 20 Elements

Element	Symbol	Atomic Number (Z)	Electron configuration
hydrogen	H	1	1
boron	B	5	2.3
oxygen	O	8	2.6
neon	Ne	10	2.8
magnesium	Mg	12	2.8.2
phosphorus	P	15	2.8.5
potassium	K	19	2.8.8.1

Another way of showing electron structure

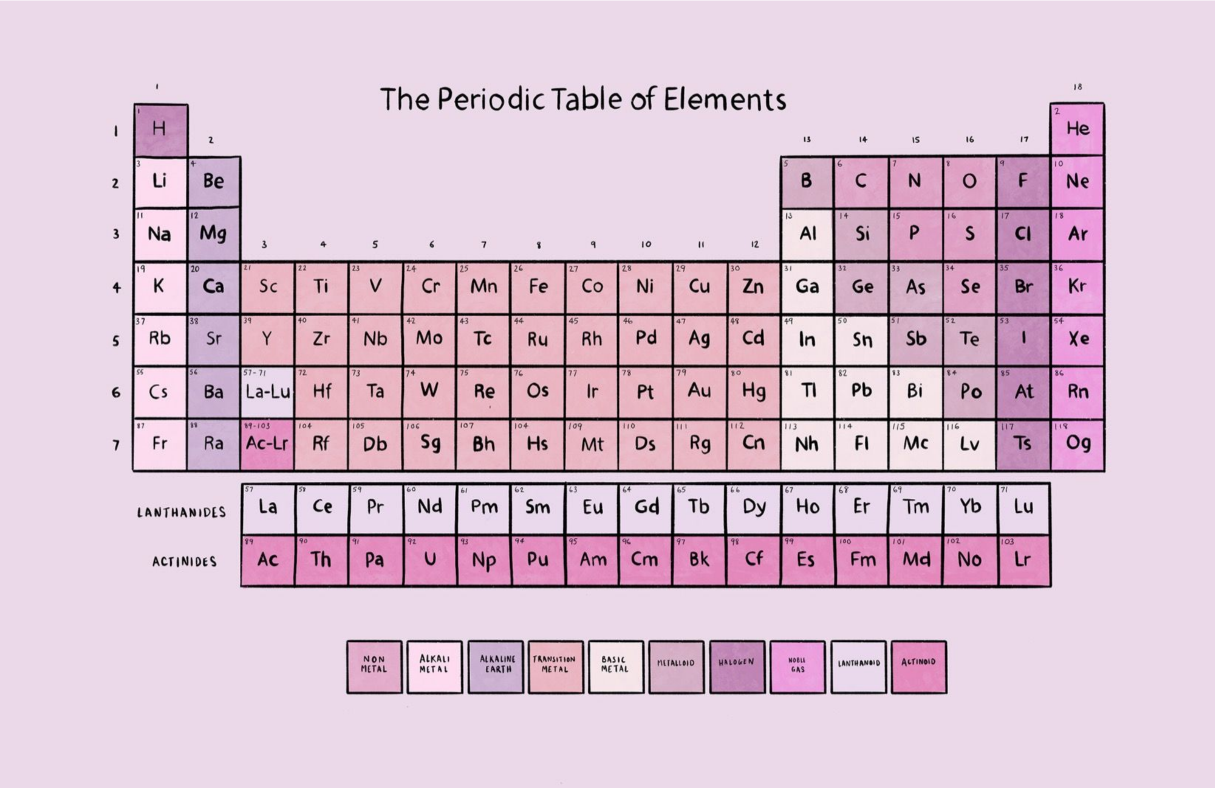






Periodic Table

- The electron configuration of an element determines the group number and period number of that element in the periodic table



Metal	**Metalloid**	**Non-metal**
Lithium	Boron	Hydrogen
Beryllium	Silicon	Helium
Sodium		Carbon
Magnesium		Nitrogen
Aluminum		Oxygen
Potassium		Fluorine
Calcium		Neon
		Phosphorus
		Sulfur
		Chlorine
		Argon

Chemical Formulae

Nomenclature of Inorganic Compounds

Binary Covalent Compounds

- Covalent / molecular compounds are form when nonmetal elements bond to each other
- The first element in the formula is simply listed using the name of the element.
- The second element is named by taking the stem of the element name and adding the suffix -ide.
- A system of numerical prefixes is used to specify the number of atoms in a molecule.

Atoms in Compound	Prefix on the Name of the Element
1	mono-*(not used for the first element's name)
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

Examples

Chemical Formula	IUPAC name
NO	nitrogen monoxide
SF_6	sulfur hexoxide
ClO_2	chlorine dioxide
NCl_3	nitrogen trifluoride
I_2O_5	diiodine pentoxide

Exceptions

- Compounds that are always called by the common names

chemical formula	common name
H_2O	water
NH_3	ammonia
CH_4	methane
PH_3	phosphine

- Common name that getting used more frequently than systematic name

chemical formula	systematic name	common name
NO	nitrogen monoxide	nitric oxide
N_2O	dinitrogen monoxide	nitrous oxide
NaCl	sodium chloride	table salt
$NaHCO_3$	sodium hydrogencarbonate/bicarbonate	baking soda
$MgSO_4$	magnesium sulfate	Epsom salt

Ionic Compounds

- Ionic compounds are formed when metal atoms lose one or more of their electrons to nonmetal atoms.
- Ionic compounds are made up of ions.
- Some elements can form two different ions, we can use two ways too distinct them:
    - **Stock System** (Modern approach)
    - Common System

Ion	stem	charge	modern name	common name
Iron	ferr-	2+	iron(II) ion	ferrous ion
		3+	iron(III) ion	ferric ion
copper	cupr-	1+	copper(I) ion	cuprous ion
		2+	copper(II) ion	cupric ion

Common Cations

Cations with 1+		Cations with 2+		Cations with 3+	
Li^+	lithium ion	Mg^{2+}	magnesium ion	Al^{3+}	aluminium ion
Na^+	sodium ion	Ca^{2+}	calcium ion	Cr^{3+}	chromium(III) ion
K^+	potassium ion	Ba^{2+}	barium ion	Fe^{3+}	iron(III) ion
Cu^+	copper(I) ion	Mn^{2+}	manganese(II) ion		
Ag^+	silver ion	Fe^{2+}	iron(II) ion		
H^+	hydrogen ion	Co^{2+}	cobalt(II) ion	**Cations with 4+**	
		Ni^{2+}	nickel(II) ion	Pb^{4+}	lead(IV) ion
		Cu^{2+}	copper(II) ion	Si^{4+}	silicon(IV) ion
		Zn^{2+}	zinc ion	Sn^{4+}	tin(IV) ion
		Sn^{2+}	tin(II) ion		
		Pb^{2+}	lead(II) ion		

Common Anions

Anions with 1-		Anions with 2-		Anions with 3-	
F^-	fluoride ion	O^{2-}	oxide ion	N^{3-}	nitride ion
Cl^-	chloride ion	S^{2-}	sulphide ion	P^{3-}	phospide ion
Br^-	bromide ion	*O_2^{2-}	peroxide ion	**Anions with 4-**	
I^-	iodide ion			C^{4-}	carbide ion
*H^-	hydride ion				

Examples

Compounds	Name
NaCl	sodium chloride
$CaBr_2$	calcium bromide
Mg_3N_2	magnesium nitride

Nomenclature of Acids

Acids that doesn't Contain Oxoanion

- Naming as Ionic compounds
  - first write the cation(usually hydrogen), then write the anion
  - Ionic name is preferred when the compound is not acting as an acid

        - i.e. pure HCl in gas phase
- Naming as Acids
    - Add prefix hydro- to the name of anion, then replace the last syllable from –ide to –ic acid.
    - Acid name is preferred when the compound acts as an acid
        - particularly when it is in solution form in water
- Examples

Compounds	**name(aqueous)**	**name(gas)**
HCl	hydrochloric acid	hydrogen chloride
HBr	hydrobromic acid	hydrogen bromide
H_2S	hydrosulfuric acid	hydrogen sulfide
HF	hydrofluoric acid	hydrogen fluoride
HI	hydroiodic acid	hydrogen iodine
*HCN	hydrocyanic acid	hydrogen cyanide

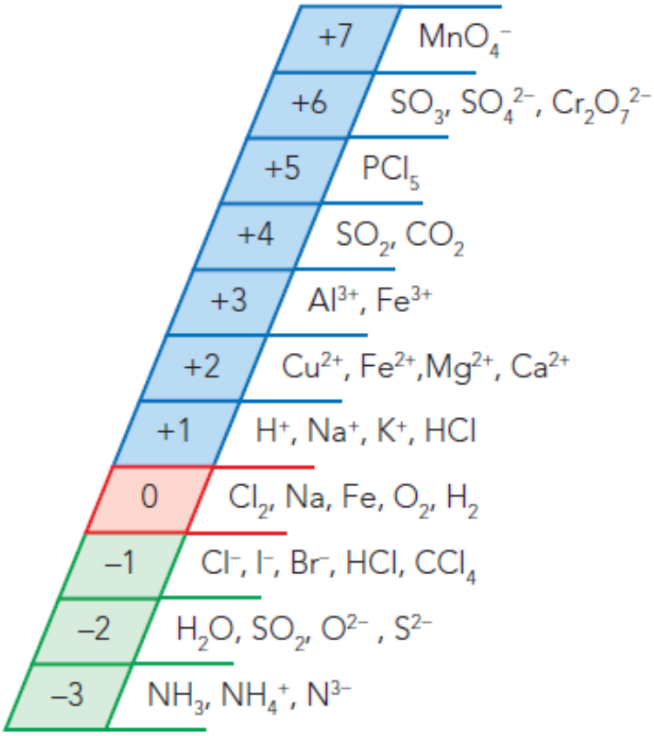
    - all above the above are binary acids, except HCN

Acids that Contains Oxoanion & Others

chemical formula	name	ion	name of ion
HNO_2	nitrous acid	NO_2^-	nitrite ion nitrate(III) ion
HNO_3	nitric acid	NO_3^-	nitrate ion nitrate(V) ion
H_2CO_3	carbonic acid	HCO_3^-	hydrogencarbonate ion bicarbonate ion
		CO_3^{2-}	carbonate(IV) ion
CH_3COOH	ethanoic acid acetic acid	CH_3COO^-	ethanoate ion acetate ion
$HMnO_4$	permaganic acid	MnO_4^-	maganate(VII) ion permaganate ion
H_2SO_3	sulfurous acid	SO_3^{2-}	sulphite ion sulphate(IV) ion
H_2SO_4	sulfuric acid	SO_4^{2-}	sulphate(VI) ion
H_2CrO_4	chromic acid	CrO_4^{2-}	chromate(VI) ion
$H_2Cr_2O_7$	dichromic acid	$Cr_2O_7^{2-}$	dichromate(VI) ion
H_3PO_4	phosphoric acid	PO_4^{3-}	phospate(V) ion
$HClO$	hydrochlorous acid chloric(I) acid	ClO^-	hydrochlorite ion chlorate(I) ion
$HClO_2$	chlorous acid chloric(III) acid	ClO_2^-	chlorite ion chlorate(III) ion
$HClO_3$	chloric acid chloric(V) acid	ClO_3^-	chlorate ion chlorate(V) ion
$HClO_4$	perchloric acid chloric(VII) acid	ClO_4^-	perchlorate ion chlorate(VII) ion
H_2O	water	OH^-	hydroxide ion
*NH_3	ammonia	NH_4^+	ammonium ion

Oxidation Numbers

- An atom's **oxidation number** (or **oxidation state**) is the **imaginary charge** that the atom would have if all of the bonds to the atom were completely **ionic**.



Guidelines

- **An atom of a free element** has an **oxidation number of 0**.
  - i.e. Cl_2, H_2, S_8, P_4, Fe, Na
- **A monatomic ion** has an oxidation number equal to its **charge**.
  - i.e. the oxidation number of Cu^{2+} is +2, Br^- is -1
- When combined with other elements
  - **alkali metals** (Group IA) always have an oxidation number of **+1**
  - **alkaline earth metals** (Group IIA) always have an oxidation number of **+2**
  - oxidation number of **aluminum ions** is always **+3**
  - **transition metal ions** have **variable oxidation number**

Transition Metal	Oxidation Number of Metal	Examples of Compounds
copper	+1	CuO_2
	+2	CuO
iron	+2	$FeCl_2$
	+3	$FeCl_3$
manganese	+2	MnO
	+4	MnO_4
	+7	Mn_2O_7

- **Fluorine** has an oxidation number of **-1** in all compounds.
- **Hydrogen** has an oxidation number of **+1** in most compounds.

- The major exception is when hydrogen is combined with metals, as in NaH, a, H or $LiAlH_4$ (**metal hydrides**). In these cases, the oxidation number of hydrogen is **-1**
- **Oxygen** has an oxidation number of **-2** in most compounds.
  - The major exception is in **peroxides** (compounds containing O_2^{2-}), where oxygen has an oxidation number of **−1**
    - i.e. H_2O_2, Na_2O_2
- The **other halogens** (Cl, Br, I) have an oxidation number of **−1** in compounds
  - unless combined with **oxygen** or **fluorine**.
    - i.e. the oxidation number of Cl n the ion ClO_4^- is +7
- **The sum of the oxidation numbers** for all atoms in a **neutral compound** is equal to **zero**
- While he sum for all atoms in a **polyatomic ion** is equal to the **charge on the ion**.
  - Consider the polyatomic ion NO_3^-. Each O atom has an oxidation number of -2. Since the overall charge on the ion is −1, the oxidation number of the N atom must be +5.
- One thing to note is that oxidation numbers are written with the **sign (+ or -) *before* the number**. This is in contrast to the charges on ions, which are written with the sign *after* the number.

Physical and Chemical Properties

Physical Properties and Physical Change

Physical Properties

- Characteristics that can be observed without changing the substances that make up the material

Examples	
boiling point	melting point
odor	lustre(the brightness that a shiny surface has)
density	attraction to magnet
electrical conductivity	thermal conductivity
color	malleability(ability to be drawn in sheets)
hardness	ductility(ability to be drawn in wires)

Physical Changes

- The substances present remain chemically the same, no new substances are formed
- often easy to reverse

Examples	Reversibility
a liquid boiling	reversible
salt dissolving in water	reversible
wood being chopped	irreversible

Examples	Reversibility
a mixture of color being separated out	reversible
peeling an orange	irreversible

Chemical Properties and Chemical Change

Chemical Properties

- Characteristics that indicate if a material can undergo a certain chemical change

Examples	
chemical stability	ability to oxide
radioactivity	half-life
flammability	acidity

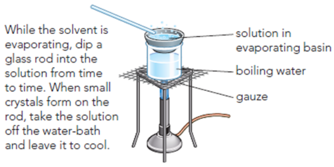
Chemical Change

- the major feature of a chemical change, or reaction, is that new substance(s) are made during the reaction
- many reactions are difficult to reverse
- during a chemical reaction, energy can be given out (exothermic change) or taken in (endothermic change)
  - there're many more exothermic change than endothermic change
- Examples
  - the formation of rust
  - baking a cake
  - photosynthesis in plants
  - explosion

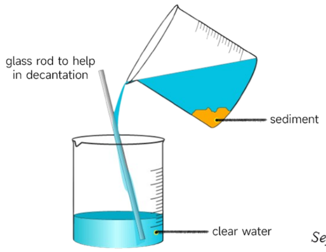
Differences between Chemical and Physical Changes

Physical Changes	Chemical Changes
don't form new substances, substances change in size, shape or state	form new substances
involves the transfer of energy from one substance to another	substances gain or lose energy, heat and light may be produced
often reversible	often irreversible

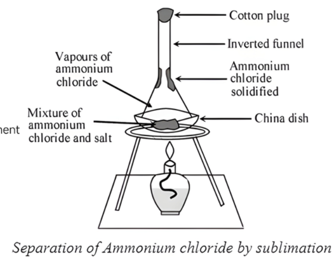
Separation of Mixtures



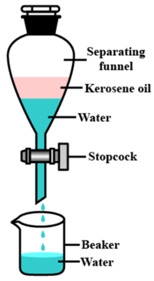
evaporation & crystallization



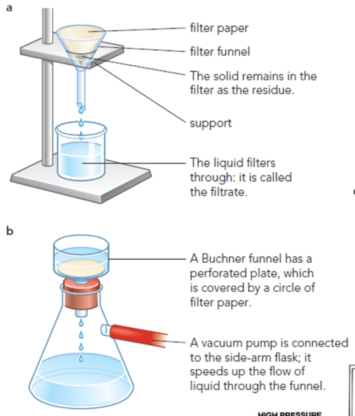
decantation



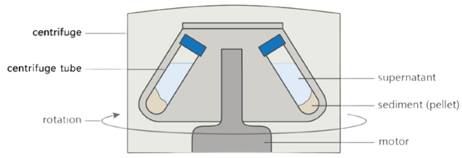
sublimation



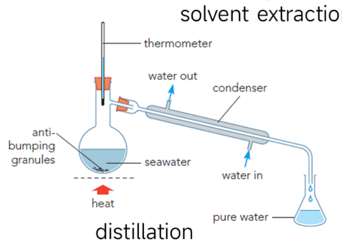
separating funnel / solvent extraction



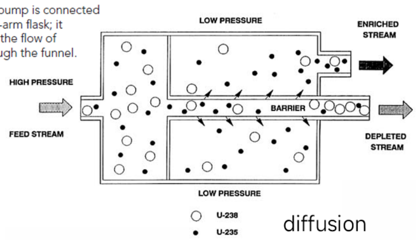
filtration



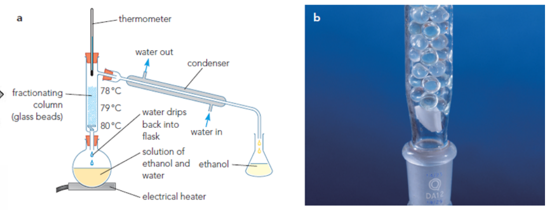
centrifugation



distillation



diffusion



fractional distillation

Details

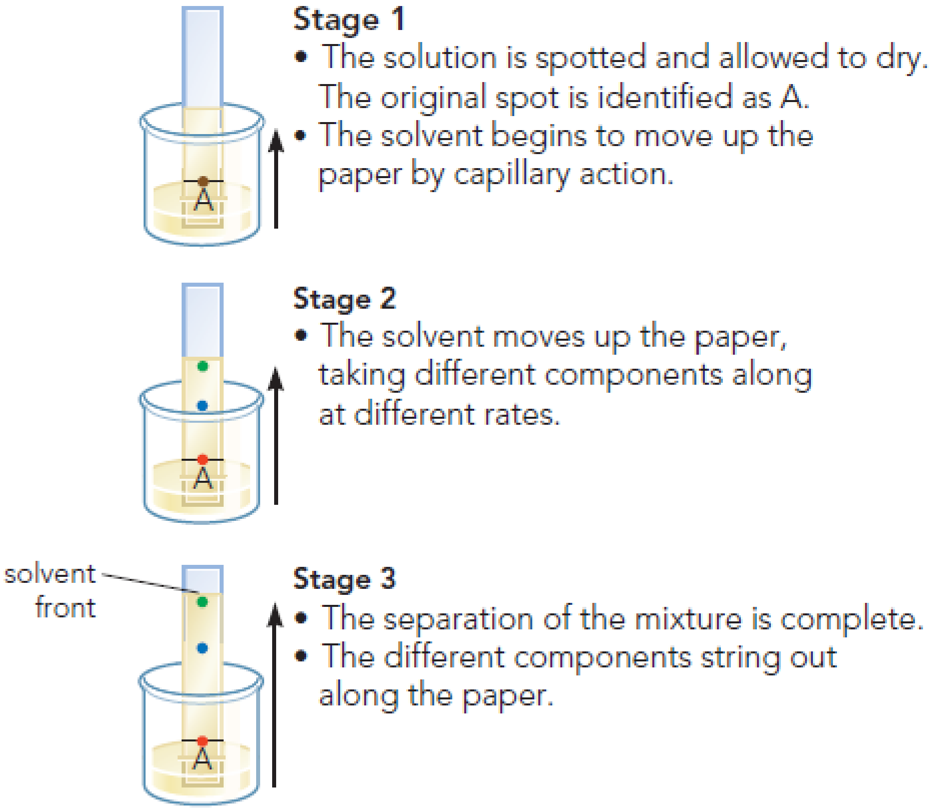
1. Evaporation
   - Used to separate a solute from a solution by **heating** the solution until the **solvent evaporates**, leaving behind a solid residue
   - *separate solid(soluble) from liquid*
2. Crystallization
   - Used to separate a solid solute from a liquid solution by **cooling the solution** until the **solute crystallizes** out.
   - The crystals can then be collected and dried to obtain a purified solid.
   - *separate solid(soluble) from liquid*
3. Filtration
   - Used to separate a solid from a liquid or gas by passing the mixture through a filter medium
   - *separate solid(insoluble) from liquid or gas*
4. Decantation
   - Used to separate a solid from a liquid by carefully pouring off the liquid without disturbing the solid at the bottom of the container
   - This method is often used in conjunction with other methods, such as centrifugation or filtration, to separate the solid from the liquid more effectively.
   - *separate solid(insoluble) from a liquid*
5. Centrifugation
   - Used to separate solid particles from a liquid by spinning the mixture at a high speed, which causes the denser particles to settle at the bottom
   - *separate solid(insoluble) from a liquid*
6. Sublimation

   - Used to separate a mixture of a solid and a volatile component that sublimes when heated.
   - Involves heating the mixture until the volatile component sublimes, leaving behind the solid residue.
   - *separate solid from a solid (one can sublimes)*
7. Distillation
   - Based on the different boiling points of the two liquids
   - Useful when the boiling points of the two liquids differ significantly
   - The mixture is heated until the liquid with the lowest boiling point evaporates, then the vapor is collected and condensed to obtain a purified liquid.
   - *separate liquid from a liquid (miscible)*
8. Separating Funnel
   - Used when the two liquids have different densities and are immiscible.
   - Involves pouring the mixture into a separatory funnel and allowing it to settle, then draining off the denser liquid from the bottom of the funnel
   - *separate liquid from a liquid (immiscible)*
9. Solvent Extraction
   - Used when the two liquids have different solubilities in a particular solvent.
   - The mixture is mixed with a solvent that selectively dissolves one of the liquids, and the two liquids are separated by draining off the solvent layer, usually using separatory funnel.
   - *separate a solid and a liquid* (dissolving the solid into a solvent)
10. Fractional distillation
   - Used when the boiling points of the two liquids are close to each other
   - Involves using a fractionating column to separate the components based on their vapor pressures
   - The mixture is heated, and the vapor is allowed to rise up through the column. As the vapor rises, it cools and condenses on the column's surface, and the more volatile component condenses and drips back into the flask. The less volatile component condenses further up the column and is collected separately.
   - *separate liquid from a liquid (miscible, similar boiling point)*
11. Diffusion
   - Used to separate the two mixtures that have different molecular sizes.
   - Diffusion is a process in which molecules move from an area of higher concentration to an area of lower concentration.
   - However, diffusion alone may not be an effective method to separate two mixtures.
   - dialysis
     - Separate the mixture through diffusion by using a semipermeable membrane.
     - A semipermeable membrane is a barrier that allows only certain molecules to pass through it.
     - By placing the mixture on one side of the membrane, the smaller molecules will diffuse through the membrane more easily than the larger molecules.
   - gas diffusion
     - The two mixtures are separated by a barrier that allows gas molecules to pass through it

- The gas molecules will diffuse through the barrier based on their size and concentration, allowing the two mixtures to be separated.

- *separate two gas*

12. Paper Chromatography



- Used to separate a mixture based on the differences in the components' solubility and attraction to a stationary phase and a mobile phase.
- *separate two or more dissolve solid in a solution*

`well I guess that's pretty much it. It isn't perfect in any way, but it took way too long time, so I guess I have to end it roughly xD.`

`jiahuiiiii @ 7th March 2023 2348`