1 Separating Mixtures

A **mixture** is made up by two or more substances mixed together, but not chemically combined.

To separate the substances we use different techniques:

A Filtration
B Evaporation
C Distillation
D Chromatographie

A Filtration

Filtration is used to separate an *insoluble* solid from a liquid. E.g to separate soil and water.

The liquid is called a **filtrate**, the mixed-in insoluble solid will we shall call the **residue**.

For filtration you need:

- *Erlenmayer flask* to receive the filtrate
- a *funnel* to keep the filter paper in place
- *filter paper*

Pour the mixture into the funnel. The liquid will pass through the filter paper and the solid substance will be left behind if the particles are too large to pass through. If not you will use a different method as we will discuss next.

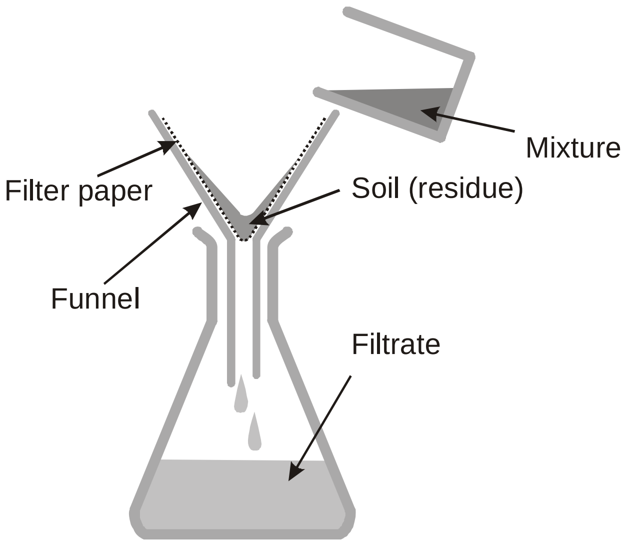


Figure 1

B Evaporation

Evaporation is used to separate an *soluble* solid from a liquid. E.g to separate salt and water.

When the mixture boils, the solvent (the water) will evaporate, leaving the solute (the salt) behind.

This method is used to retain the solid residue. If we want to retain the vapour we must use distillation as is shown in the next experiment.

What you need:

- *Tripode*
- *Wire gauze*
- *Bunsen burner*
- *Beaker* for the solution

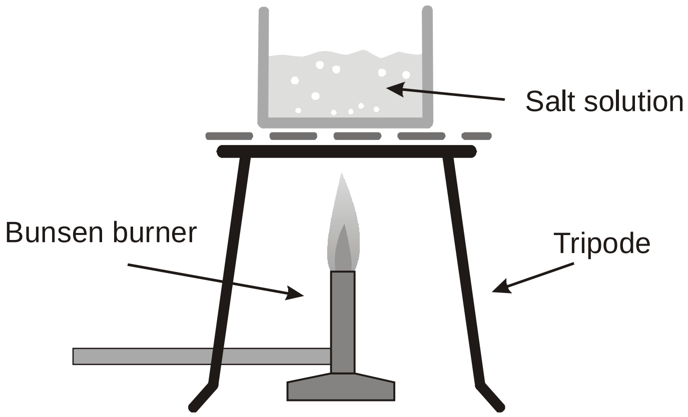


Figure 2

C Distillation

Distillation is used to separate to liquids with different boiling points. E.g to separate water and alcohol.

Distillation can also be used to separate salt and a liquid if you don't want to loose the liquid.

What you need:

- *Tripode*
- *Wire gauze*
- *Bunsen burner*
- *round-bottomed flask*
- *Thermometer*
- *Stopper with one tube*
- *Liebig condenser*
- *Beaker* for the distillate

If the solution is heated up to the boiling point of the lower boiling liquid, this liquid will evaporate and it's vapour can be cooled down and condensed in the Liebig condenser and the reformed liquid will be caught in the beaker as shown in figure 3.

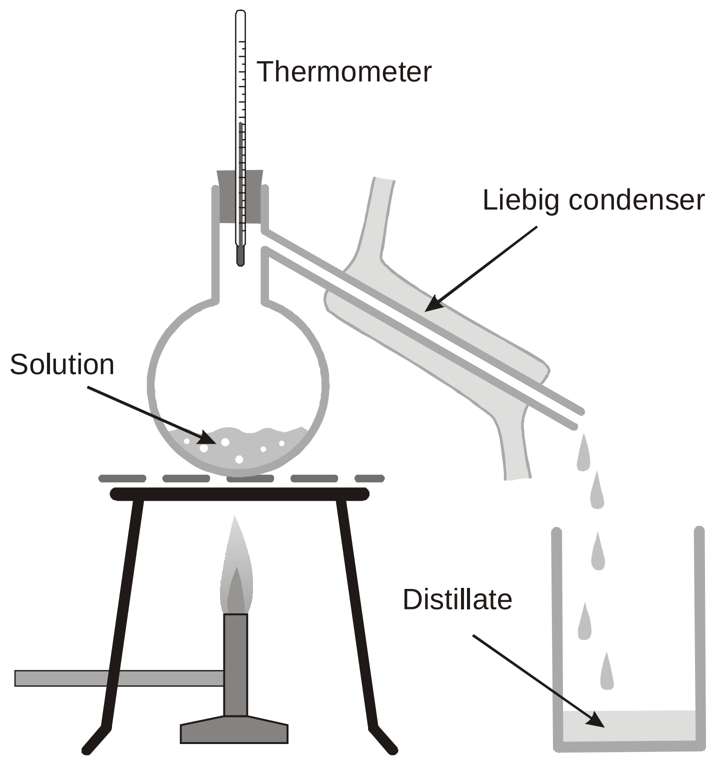


Figure 3

D Chromatographie

Chromatographie can be used to separate a mixture of dissolved substances in a solution.
E.g. a mixture of the different components of the colours of a marker.

What you need:

- *Graduated cylinder*
- *Solvent (water, alcohol)*
- *Chromatography paper*

Make two ink spots from different markers on the chromatography paper. Place the lower end of the chromatography paper into the solvent, just beneath the ink spots.

As the solvent travels up the chromatography paper, some dyes are carried further than others and hence are separated due to their differing chemical structures.

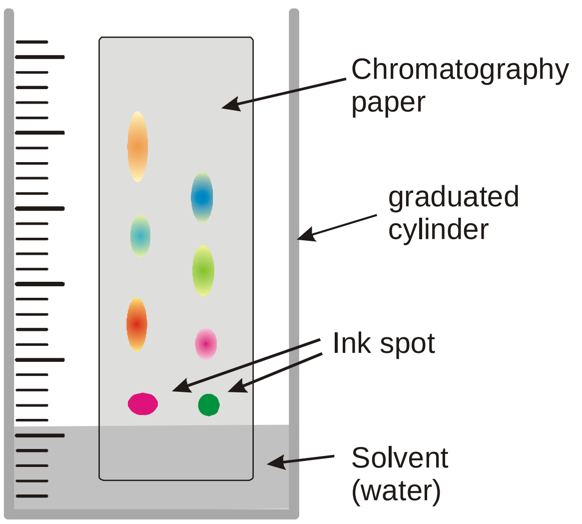


Figure 4

Key Points
• A mixture is made up of two or more substances mixed together but not chemically combined.
• Filtration can be used to separate a liquid and an insoluble solid.
• Evaporation is a method used to separate soluble solids from liquids by boiling off the liquid.
• Distillation separates mixed liquids by vaporising and condensing the liquid with the lower boiling point.
• Distillation can also be used to separate a soluble solid from a liquid.
• Chromatography is a technique used to separate a mixture of dissolved substances in a solution.

2 States of Matter

Matter is anything that occupies space and has mass.

Examples: A piece of copper is matter.
Water is matter.

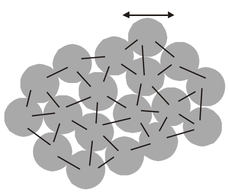
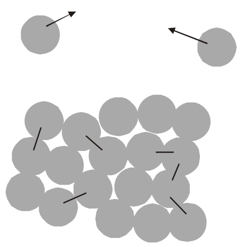
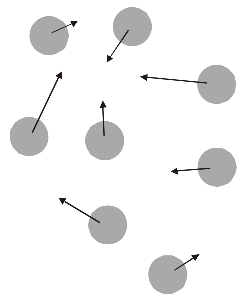
But: Terms like *luck, love, jealousy, patience* etc. are not matter, they don't occupy space, they don't have mass.

Matter is made up of particles: Atoms or molecules.

As you know from water, matter exists in three states:

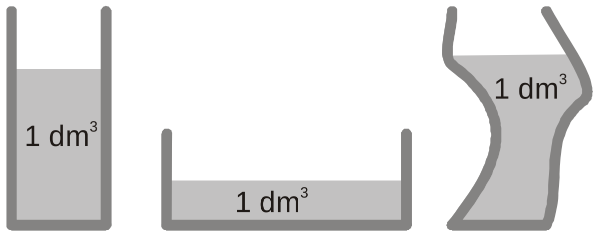
- Solid
- Liquid
- Gas

If matter is solid, liquid or gaseous, the particles the matter is made of are the same. They just behave differently.

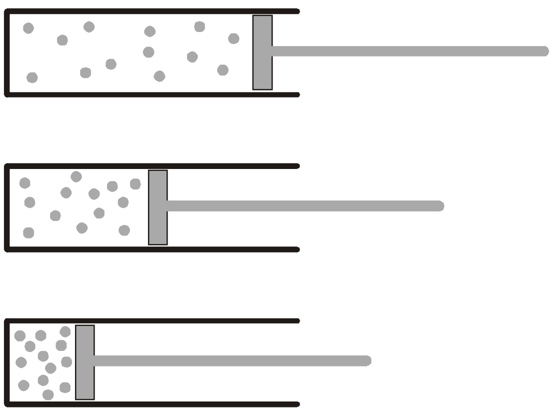
Solids	**Liquids**	**Gases**
		
In **solids** the particles are in fixed position. They can vibrate but cannot move past each other.	In **liquids** the particles can slide past each other, but cannot split from the group (some do, they evaporate).	In **gases** the particles move in any direction independent of each other (like a swarm of mosquitoes).
Definite shape	No definite shape	No definite shape
Do not take the shape of their container	Take the shape of their container	Take the shape of their container
Have a definite volume	Have a definite volume	No definite volume
Cannot be compressed	Cannot be compressed	Can be compressed or decompressed

Give an example of matter in a solid, liquid and gaseous form.

A liquid takes the shape of its container. One kilogram of water takes a volume of one litre equal to one decimetre cubed. Its shape can be different, depending on the container.

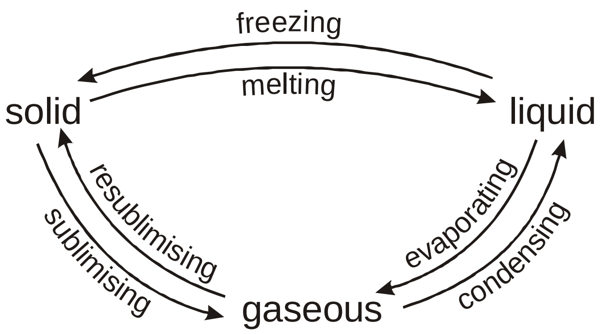


A gas doesn't have a fixed volume. It can be compressed or decompressed.



Matter can be changed from one state to another by either heating or cooling.

Heating makes ice melt.
Heating makes water evaporate.
Cooling makes vapour condense.
Cooling makes water freeze.



About the weather

The change in form when it rains is called

……………………………………………………..

The change in form when it snows is called

……………………………………………………..

The change in form when it is foggy is called

……………………………………………………..

The change in form when there are hailstones

……………………………………………………..

Key Points
• Matter is anything that occupies space and has mass. • The three states of matter are solid, liquid and gas. • Substances can be changed from one stage to another by either heating or cooling them.

3 Classification of Substances

All substances are made up of **atoms**[1].
Atoms cannot be divided by means of chemistry.

Substances, that are made up by only one type of atom, are called **elements.**
In nature there are 92 different elements[2].

Some important elements:

H	Hydrogen	K	Potassium
He	Helium	Na	Sodium
C	Carbon	Fe	Iron
N	Nitrogen	Cu	Copper
O	Oxygen	Al	Aluminium
S	Sulphur	Au	Gold
Cl	Chlorine	U	Uranium

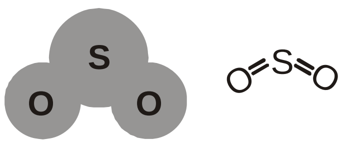
Most of the substances are made up of **molecules.**
Molecules often are called **compounds.**
A molecule is a combination of at least two atoms that are held together by forces called *chemical bounds.*

Examples:

Each molecule of **water** consists of two atoms of Hydrogen and one atom of Oxygen.

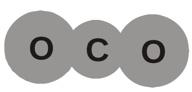
H_2O

Each molecule of **sulphur dioxide** consists of two atoms of Hydrogen and one atom of Oxygen.



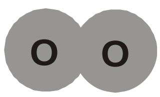
SO_2

Each molecule of **Carbon dioxide** consists of two atoms of Carbon and one atom of Oxygen.



O=C=O CO_2

Oxygen as found in the atmosphere consists of two atoms of oxygen bound together.



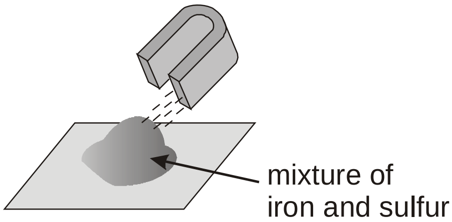
O=O O_2

[1] The Greek word ατομοσ means *indivisible*.
[2] Some more are produced artificially, but they are not stable.

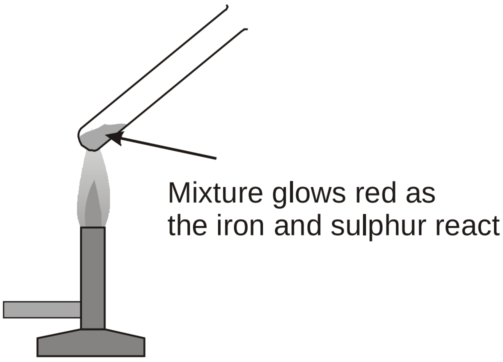
To compare the properties of a compound to its elements

Iron is a grey metal. It is attracted to a magnet. Sulphur is a yellow solid.

You can mix *iron dust* with *sulphur dust*. Iron is still attracted to a magnet.



If the mixture of *iron dust* (Fe) and *sulphur dust* (S) is heated, iron and sulphur will *react* with each other. The result is *iron sulphide* (FeS).



In this reaction heat is deliberated. The Bunsen burner just initiates the reaction.

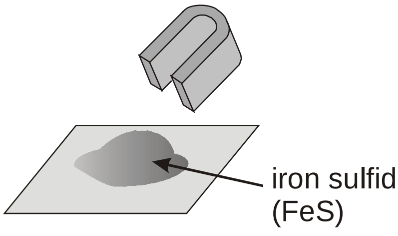
The reaction can be represented by a chemical equation:

$$Fe + S \rightarrow FeS + \text{heat (energy)}$$

To make clear, that all substances that are involved are *solids*, we add an (s) behind.

$$Fe(s) + S(s) \rightarrow FeS(s) + \text{heat (energy)}$$

The properties of iron sulphide are different to the properties of the two substances it is made of. So the iron sulphide is no longer attracted to a magnet.



Combustion of Carbon

Burning **carbon** (coal, wood) means a reaction of carbon with **oxygen.** Carbon is a solid, Oxygen is a gas.

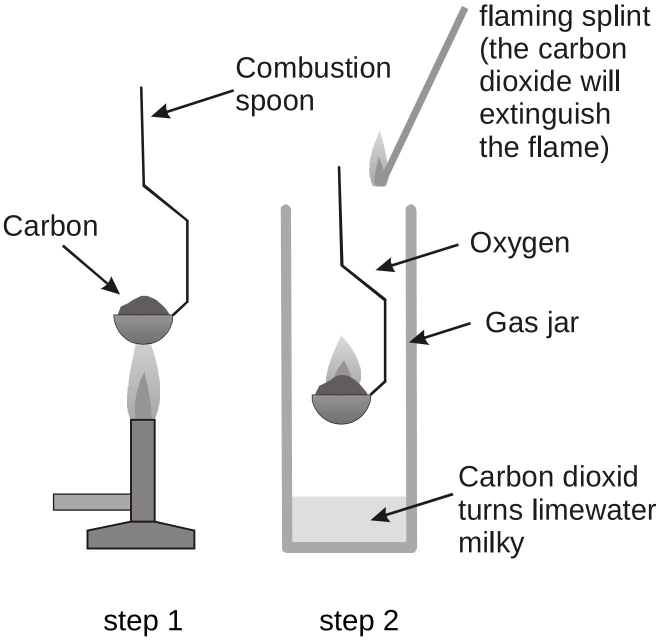
$$C(s) + O_2(g) \rightarrow CO_2(g) + \text{heat (energy)}$$

Carbon and oxygen burn to form carbon dioxide.

For burning carbon you use:

- Bunsen burner
- Combustion spoon
- Gas jar
- Lime water

If you burn carbon within a gas jar, carbon dioxide is produced. Carbon dioxide is a gas.



Oxygen (O_2) is a gas that is needed to <u>burn</u> substances such as carbon, wood, petrol etc.

Carbon dioxide (CO_2) is a gas also. It can be used to <u>extinguish</u> fires.

Carbon dioxide turns limewater milky.

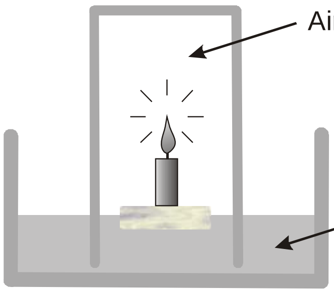
Carbon dioxide extinguishes fires.

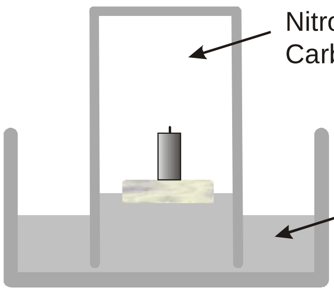
A candle burning inside a glass jar

What you need:

- Container
- Glass jar
- Piece of wood
- Candle

Fill a container with water, put a piece of wood into the water so that it floats. Put a lit candle on it and place a beaker over the candle. Watch what happens.





Wax consists of big molecules made up of carbon and hydroxide; wax is a **hydrocarbon.** As the wax is heated, it melts and climbs up the wick. There it will be broken into smaller molecules and will react with the oxygen (O_2).

This reaction will cause hydrogen to produce water molecules:

$$4H(s) + O_2(g) \rightarrow 2H_2O(g) + \text{energy}$$

The H_2O will condense to its liquid form.

The carbon will glow and giving light and it will react with oxygen to form carbon dioxide:

$$C(s) + O_2(g) \rightarrow CO_2(g) + \text{energy}$$

As the level of oxygen decreases, the flame will burn out due to the lack of oxygen.

As the build up water vapour condenses, the carbon dioxide will partly dissolve in water (as you know from the bubbling water), hence the volume inside the glass jar decreases and the level of water inside the glass jar will rise.

Questions

1. What is an element. Give two examples.

2. Distinguish between a compound and a mixture

Key Points
• Elements are made up of only one type of atom. • Molecules consist of two or more atoms chemically combined. • A mixture contains two or more different substances mixed together but not chemically combined. • The properties of a compound differ from those of the elements from which it is made.

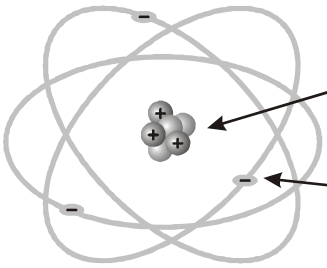
4 Atomic Structure

Atoms are the smallest parts of an element. By means of chemistry they are not further divisible. But they can be divided into smaller parts by means of atomic physics.

Atoms consist of three sub-atomic particles:

- **Protons**
- **Electrons**
- **Neutrons.**

The nucleus is made up of protons and neutrons. It is surrounded by electrons.



The electrical charge of a proton is called +1.
The mass of a proton is 1, it represents the *atomic mass unit*.

A neutron is electrically neutral, its mass is equal to the mass of a proton (even a little bit higher), so the mass of a neutron is also 1.

An electron is negatively charged to an amount of –1.
Its mass is about one two-thousandth $\left(\frac{1}{2000}\right)$ the mass of a proton or a neutron, roughly we can say the mass of an electron is zero.

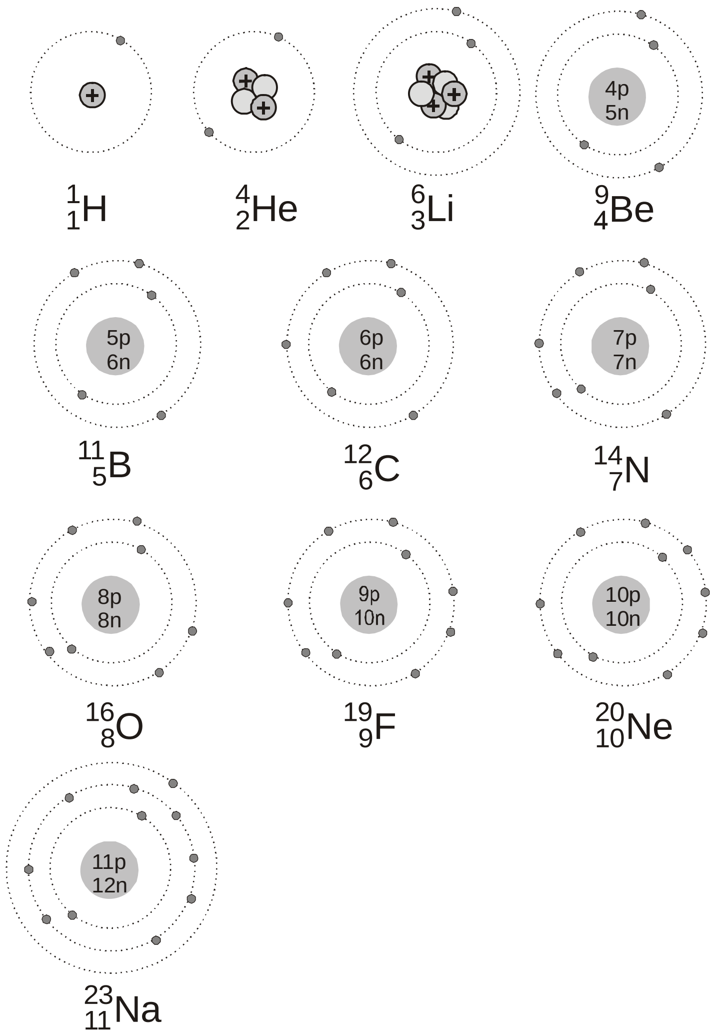
In an atom the number of protons is equal to the number of electrons, hence an atom is electrically neutral.

Electrons are arranged in distinct shells (see diagram). The innermost shell can contain two electrons, not more. The third electron has to take its place in the next shell, which can contain 8 electrons in total.

From the diagram we can see the structure (the *electron configuration*) of the elements hydrogen (H), lithium (Li), beryllium (Be), boron (B), carbon (C), nitrogen (N), oxygen (O), fluorine (F), neon (Ne) and sodium (Na).

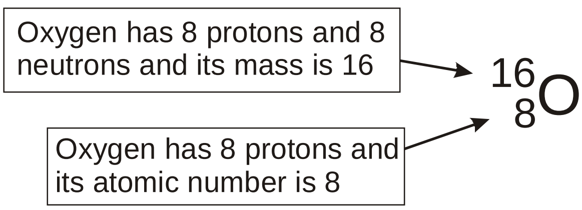
With Neon the second shell is complete, it has 8 electrons in its outer shell.

Now we look at sodium. It has 11 protons, 12 neutrons and 12 electrons. The first shell is full with 2 electrons and the second shell is full with 8 electrons, therefore the 11th electron has to take its place on the third shell.



The meaning of the numbers above and below the symbol of an element are:

- The number below gives the amount of protons (and also the *Atomic Number*).
  This number also gives the number of electrons.
- The number above gives the amount of protons <u>and</u> neutrons (and also its *mass*).



Maybe you have realised that the number of neutrons is roughly equal to the number of protons.

$^{1}_{1}H$ (Hydrogen) has one proton but no neutrons.

$^{4}_{2}He$ (Helium) has two protons and two neutrons.

$^{16}_{8}O$ (Oxygen) has eight protons and eight neutrons.

$^{23}_{11}Na$ (Sodium) has 11 protons and 12 neutrons.

What is the function of neutrons?

Equally charged particles repel each. A nucleus could never exist of only protons, because the repelling forces would cause the nucleus to explode.

Neutrons serve as a sort of glue. They stabilise the nucleus.

Hydrogen has no need for a neutron. One proton is stable on its own.

As the number of protons in the nucleus increases, the ratio of neutrons to protons must also increase.

$^{63}_{29}Cu$ (Copper) has 29 protons and 34 neutrons.

$^{238}_{92}U$ (Uranium) has 92 protons and 146 neutrons.

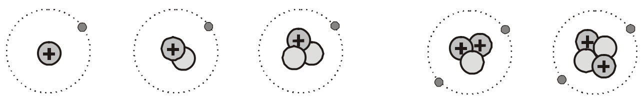
I'm sure you have already realised, that the amount of neutrons can be calculated by subtracting the number of protons (the atomic number) from the number of protons and neutrons (the mass number).

Example Uranium: 238 – 92 = 146.

About isotopes

Whether an element is classified as nitrogen, or oxygen, depends on the amount of protons in its nucleus. Nitrogen has to have five protons, otherwise it's not nitrogen. oxygen has to have six protons, otherwise it's not oxygen[3].

The amount of neutrons may differ.



$^{1}_{1}H$ $^{2}_{1}H$ $^{3}_{1}H$ $^{3}_{2}He$ $^{4}_{2}He$

Hydrogen can have none, one ore two neutrons. Normally it has none.

Helium can have one or two neutrons. Normally it has two.

Copper exists as $^{63}_{29}Cu$ with 29 protons and 34 neutrons, but also as $^{65}_{29}Cu$ with 29 protons and 36 neutrons.

Uranium exists as $^{238}_{92}U$ (U-238) and as $^{235}_{92}U$.(U-235).

U-238 is stable, U-235 is not. U-235 is used for nuclear reactors and – unfortunately – for atomic bombs.

Two elements with an equal amount of protons and a different amount of neutrons are called **isotopes.**

[3] An „oxygen-atom" with seven protons wouldn't be an oxygen-atom, it would be a fluorine-atom.

Questions

1. A proton is a sub-atomic particle. Name the other two particles and give their location, mass and charge.

Particle	***Location***	***Mass***	***Charge***
Proton			

2. What information is given by:
   a) The atomic number of an element?

   b) The mass number of an element?

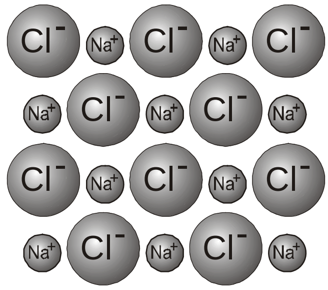
3. What do you get by

   mass number minus *atomic number* = ...

4. Draw an atomic diagram of Ne.

Key Points
• An atom is made up of three sub-atomic particles – protons, electrons and neutrons. • The atomic number is the number of protons. • The mass number is the number of protons and neutrons. • The first shell can hold two electrons and the second and third can hold up to eight each. • Isotopes are atoms which have the same number of protons but a different number of neutrons.

5 Chemical Bonding



Why elements react

All atoms want to have full outer shells of electrons. They achieve this by bonding.

Two ways of achieving a full outer shell

1. One atom looses and another atom gains electrons. This is called **ionic bonding.**
2. Two atoms overlap and share electrons. This is called **covalent bonding.**

Ionic bonding

A full outermost shell contains 8 electrons.

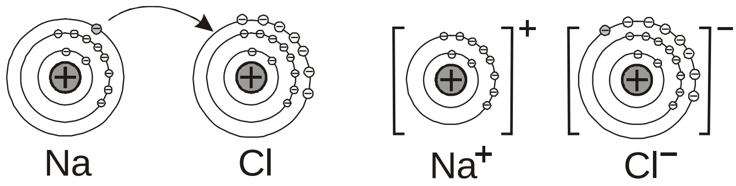
Non-metals have nearly full outermost shells. They try to gain electrons to complete their outermost shell.

Metals have one, two or three electrons in their outermost shell. They try to get rid of them.

Lets take *table salt*. It's chemistry term is *sodium chloride.*

If A wants to get rid of a bicycle and B needs a bicycle then A will give his bicycle to B and both are happy.

Sodium wants to get rid of its single electron in its outermost shell and Chlorine needs one electron to complete its outmost shell. So sodium chloride gives its electron to chlorine and they both are happy.



Sodium and Chlorine are transformed into **ions**.

Na^+ is called *sodium ion*.

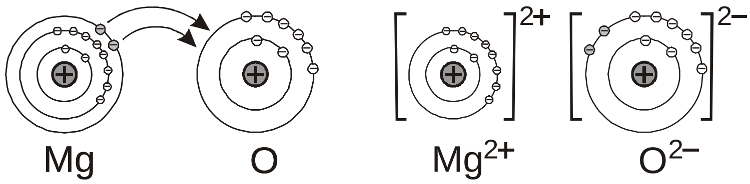
Cl^- is called *chloride*.

Na^+Cl^- (simply NaCl) is called *sodium chloride*.

The positive and the negative charges of the sodium ion and of the chloride ion attract each other. Table salt (NaCl) is a solid.

NaCl is not a single molecule. Its structure is that of a crystal.

An other example is magnesium oxide (MgO).
Magnesium donates its two electrons in its outermost shell and gives it to oxygen. You also might say, Oxygen “robs” two electron from the magnesium atom. Oxygen is known for its strong attraction to electrons.



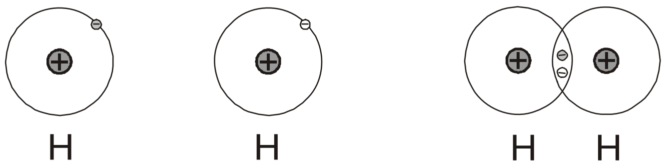
Covalent bonding

Atoms share electrons with each other to have a full outer shell. Let’s look at a few examples:

Hydrogen H_2

Each hydrogen atom has on electron in its shell. Two hydrogen atoms share their single electrons with each other by overlapping their outer shells.

It is written as **H – H** or just $\mathbf{H_2}$.



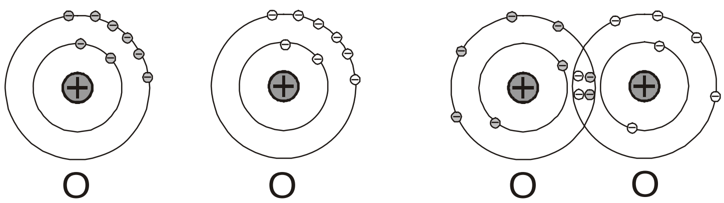
The two hydrogen atoms are *singularly-bound.*

Oxygen O_2

Each oxygen atom has six electrons in its outer shell. Two atoms of oxygen share two pairs of electrons with each other, resulting in two covalent bonds.

It is written as **O = O** or just $\mathbf{O_2}$.

The two oxygen atoms are *doubly-bounded.*



Nitrogen N_2

Each nitrogen atom has five electrons in its outer shell. Two nitrogen atoms share three pairs of electrons with each other, resulting in three covalent bonds.

It is written as **N ≡ N** or just $\mathbf{N_2}$.

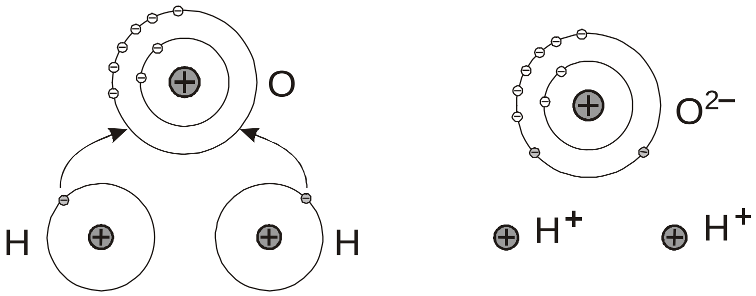
The two nitrogen atoms are bound tight to each other and it is difficult to separate them.

The two nitrogen atoms are *triple-bounded.*

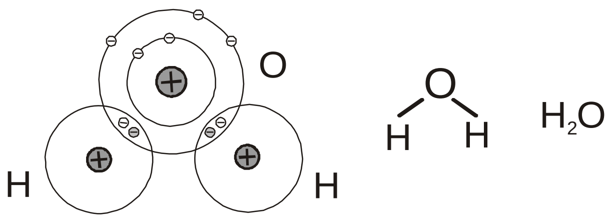
Water H_2O

A water atom consists of oxygen and hydrogen.

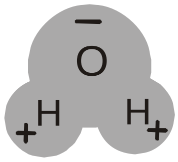
If hydrogen were a metal, the bonding could be ionic. The oxygen atom takes one electron from each of the two hydrogen atoms, forming the oxygen ion O^{2-}, leaving behind two hydrogen ions H^+ as shown beneath.



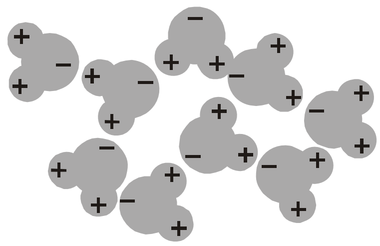
But it is not like this. Hydrogen and Oxygen have “decided” to share some electrons to build a molecule named H_2O.



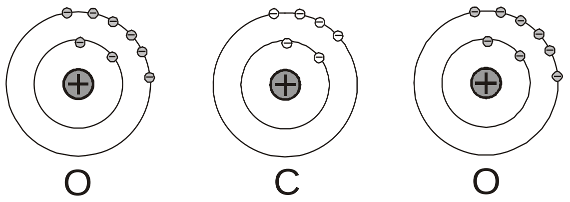
In fact, oxygen is a bit stronger ad attracting electrons. So the water molecule is slightly negatively charged at the side of the oxygen atom and slightly positively charged at the side of the two hydrogen atoms. The water molecule is called a **polarised molecule,** it has **polarity.**

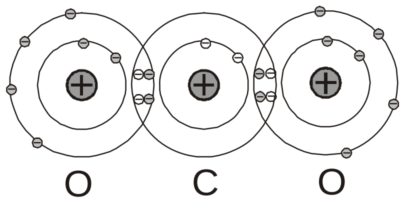


Considering this you can understand the way water molecules are tightened to each other when forming water or ice. Each positively charged side tries to be connected with a negatively charged side of another molecule. In the case of ice the connections are fixed. In the case of water they are variable and water can take the shape of its container.



Carbon dioxide CO_2

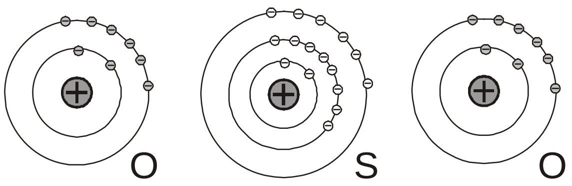


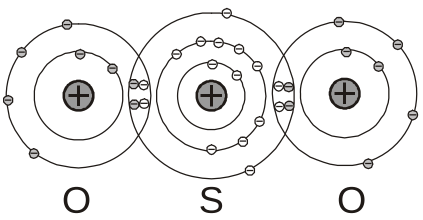


O=C=O

CO_2

Sulphur dioxide SO_2

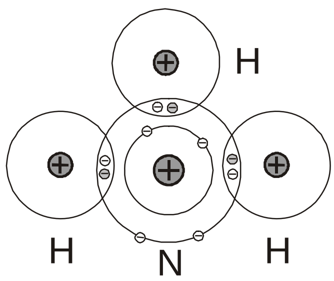


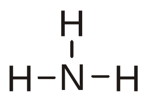


O=S=O

SO_2

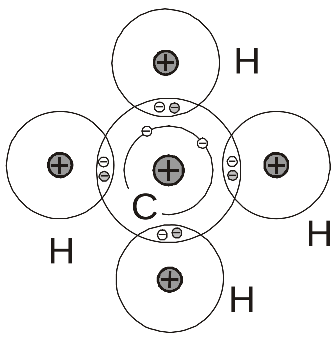
Ammonia NH_3

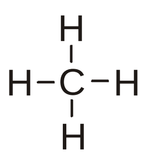




NH_3

Methane CH_4





CH_4

Properties of ionic and covalent compounds

Ionic compounds

- usually solids
- high melting and boiling temperatures
- soluble in water
- conduct electricity when dissolved in water

Covalent compounds

- usually liquids or gases
- low melting and boiling temperatures
- insoluble in water
- they are not usually good conductors of electricity

Questions

1. Why do elements react with each other?

2. What is an ionic bond?

3. What is a covalent bond? Give an example.

Key Points

- Elements react to achieve a full outer shell of electrons.
- An ion is an atom that has lost or gained electrons to complete its outermost shell.
- An ionic bond is the electrical force of attraction between positive and negative ions.
- A covalent bond consists of a pair of electrons shared between two atoms.

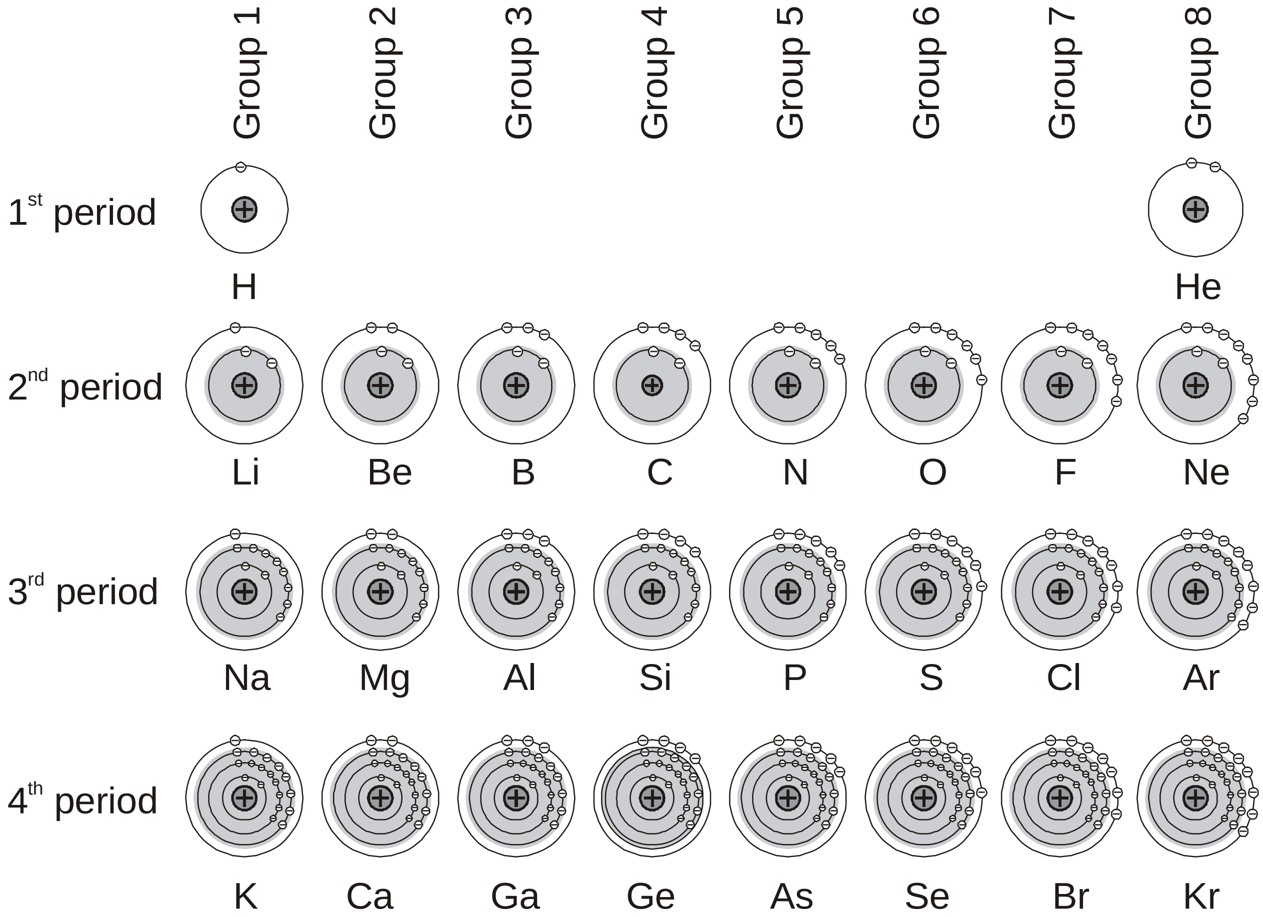
6 The Periodic Table

Groups → **Periods ↓**	Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8
1st period	1 **H** 1 Hydrogen							41 **He** 2 Helium
2nd period	6 **Li** 3 Lithium	9 **Be** 4 Beryllium	11 **B** 5 Boron	12 **C** 6 Carbon	14 **N** 7 Nitrogen	16 **O** 8 Oxygen	19 **F** 9 Fluorine	20 **Ne** 10 Neon
3rd period	23 **Na** 11 Sodium	24 **Mg** 12 Magnesium	27 **Al** 13 Aluminum	28 **Si** 14 Silicon	31 **P** 15 Phosphorus	32 **S** 16 Sulphur	35 **Cl** 17 Chlorine	40 **Ar** 18 Argon
4th period	39 **K** 19 Potassium	40 **Ca** 20 Calcium	70 **Ga** 31 Gallium	73 **Ge** 32 Germanium	75 **As** 33 Arsenic	79 **Se** 34 Selenium	80 **Br** 35 Bromine	84 **Kr** 36 Krypton
5th period	85 **Rb** 37 Rubidium	88 **Sr** 38 Strontium	115 **In** 49 Indium	119 **Sn** 50 Tin	122 **Sb** 51 Antimony	128 **Te** 52 Tellurium	127 **I** 53 Iodine	131 **Xe** 54 Hydrogen
6th period	133 **Cs** 55 Cesium	137 **Ba** 56 Barium	204 **Tl** 81 Thallium	207 **Pb** 82 Lead	209 **Bi** 83 Bismuth	209 **Po** 84 Polonium	210 **At** 85 Astatine	222 **Rn** 86 Radon
7th period	223 **Fr** 87 Francium	226 **Ra** 881 Radium						

In 1869 the Russian chemist *Dimitri Mendeleev* saw that if the elements were arranged in order of increasing atomic mass, certain properties recurred in regular intervals. But he found holes in his table because some elements had not yet been discovered.

In fact, his periodic table was not widely accepted by chemists until the missing elements were discovered and shown to have the properties that Mendeleev predicted.

Groups and Periods



Groups

The diagram beneath shows the **electron configuration** of the first 26 elements.

Each vertical column in the periodic table is called a **group.** Elements of the same group have the same number of electrons in their outermost shells and therefore the same (or similar) chemical properties.

Group 1 consist of the so called *alkali metals.* They all have one electron in their outermost shell. They all react heavily with water.

Group 2 consist of the *alkaline earth metals.* They all have two electrons in their outermost shell.

Group 7 consist of the *halogens.* They all have seven electrons in their outermost shell.

Group 8 are the *noble gases.* They all have eight electrons in their outermost shell. They don’t react at all.

Periods

Each row of elements in the table is called a **period.**

The 1st period contains of only two elements: Hydrogen and Helium. The reason is, that on the innermost shell there is only place for two electrons.

The third to sixth period contains eight elements each. The reason is, that their outermost shell contains up to 8 electrons.

Valency

The valency of an element is the number of electrons an atom wants to gain, loose or share in order to achieve a full outer shell.

Element	**needs to …**	**Valency**
Hydrogen	loose or share 1 electron	1
Group 1	loose 1 electron	1
Group 2	loose 2 electrons	2
Group 3	loose 3 electrons	3
Group 4	loose or gain or share 4 electrons	4
Group 5	gain or share 3 electrons	3
Group 6	gain or share 2 electrons	2
Group 7	gain or share 1 electron	1
Group 8	*the noble gases*	0

Metals – Nonmetals – Metalloids

Groups→ Periods ↓	1	2	3	4	5	6	7	8
1	H							He
2	Li	Be	B	C	N	O	F	Ne
3	Na	Mg	Al	Si	P	S	Cl	Ar
4	K	Ca	Ga	Ge	As	Se	Br	Kr
5	Rb	Sr	In	Sn	Sb	Te	I	Xe
6	Cs	Ba	Tl	Pb	Bi	Po	At	Rn
7	Fr	Ra						

The elements on the left of the periodic table are *metals*. The elements on the right are *non-metals*. In between some elements which show metallic properties as well as non-metallic properties. They are called *metalloids*.

Metals have one, two or three electrons in their outermost shell. Group 1 (the alkali metals) and group 2 (the alkaline earth metals) are typical metals. Iron, copper, silver, gold etc., all belong to the metals.

Non-metals have five, six, seven or eight electrons in their outermost shell. Typical non-metals are the elements of group 7 (halogens) and group 8 (noble gases).

The metallic properties decrease from the left to the right while the non-metallic properties increase. The metallic properties also increases from top to bottom while the non-metallic properties decrease.

Hydrogen is an exception. It can’t be considered a metal and it can’t be considered a non-metal and it is not a metalloid either.

Typical properties of metals

Metals are

- good electrical conductors
- good heat conductors
- shiny
- ductile and malleable (change in form by hammering)
- hard and dense
- have a high melting point

Typical properties of non-metals

Nonmetals are

- poor electrical conductors
- poor heat conductors
- brittle
- have a low melting point

Questions

1. What is the periodic table and how is it arranged?

2. What is the name of the group 1 elements? List two elements of this group.

3. What is the name of the group 7 elements?

4. What group is known as the alkaline earth metals?

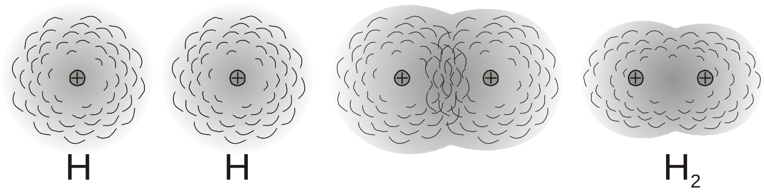
5. What is meant by the valency of an element?

Key Points
• The periodic table is an organized arrangement of the elements. • Vertical columns are called groups. • Horizontal rows are called periods. • Elements of the same group are chemically similar. • The Elements to the left are metals. • The Elements to the right are non-metals. • The Elements in between are metalloids. • The valency of an element is the number of electrons an atom wants to gain, loose or share in order to achieve a full outer shell.

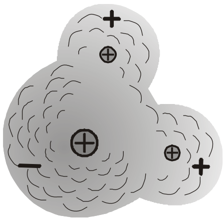
7 Electronegativity

Until now we considered electrons to be tiny particles orbiting around the nucleus. This idea is quite useful, but not quite correct. More correct is it to think electrons are surrounding the nucleus like a *cloud*.

If two atoms are bond together in a covalent bond, their electron clouds overlap building up a new cloud, called a covalent bond.



In the case of water, the electron clouds overlap in a similar way. Oxygen has a stronger force of electron attraction than hydrogen has, so the water atom is **polarized**. The side of oxygen is slightly positively charged, while the side of the hydrogen is slightly negatively charged.



When a lump of **sugar** dissolves in water it is split up into smaller pieces, it is broken down into individual molecules.

sugar (solid) $\rightarrow$ sugar (solution)

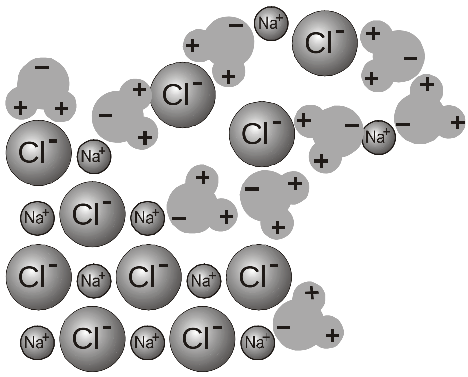
Sugar ***dissolves*** in water.

When a lump of **salt** (sodium chloride) dissolves in water, it also split up into smaller pieces. However, salt will be completely split up into its individual ions.

This is because of the **polarisation** of the water molecule. It weakens the attracting forces between the charged ions.

NaCl(solid) $\rightarrow$ $Na^+ + Cl^-$

This is called *dissociation,* the ions ***dissociate,*** they are split up into ions.



Electronegativity

Electronegativity is a quantitative measure of an atom's tendency to attract electrons in a chemical bond.

The larger the electronegativity, the greater the tendency of an atom to pull electrons towards itself in a bond.

The chart shows the electronegativity of the most important elements.

	1	2											3	4	5	6	7	8
1	**H** 2.2																	**He** 0
2	**Li** 1.0	**Be** 1.5											**B** 2.0	**C** 2.5	**N** 3.0	**O** 3.5	**F** 4.0	**Ne** 0
3	**Na** 0.9	**Mg** 1.2											**Al** 1.5	**Si** 1.8	**P** 2.1	**S** 2.5	**Cl** 3.0	**Ar** 0
4	**K** 0.8	**Ca** 1.0						**Fe** 1.8	**Co** 1.9	**Ni** 1.9	**Cu** 1.9	**Zn** 1.6	**Ga** 1.6	**Ge** 1.8	**As** 2.0	**Se** 2.4	**Br** 2.8	**Kr** 0
5	**Rb** 0.8	**Sr** 1.0									**Ag** 1.9	**Cd** 1.7	**In** 1.7	**Sn** 1.8	**Sb** 1.9	**Te** 2.1	**I** 2.5	**Xe** 0
6	**Cs** 0.7	**Ba** 0.9								**Pt** 2.2	**Au** 2.4	**Hg** 1.9	**Tl** 1.8	**Pb** 1.9	**Bi** 1.9	**Po** 2.0	**At** 2.2	**Rn** 0
7	**Fr** 0.7	**Ra** 0.9																

This chart of electronegativity allows us to explain and to predict the behaviour of chemical substances.

Oxygen has a electronegativity of 3.5, hydrogen only 2.2. Oxygen is much stronger at attracting electrons than hydrogen. This explains why the molecule H_2O is polarized.

Pure water itself can dissociate. But only to a very small amount. If one water molecule bumps into another with enough energy, a *hydrogen ion* (H^+) can split up, leaving behind a *hydroxide ion* ($[OH]^-$).

$$H_2O \rightarrow H^+ + [OH]^-$$

The reaction also goes the other way round.
H^+ and $[OH]^-$ will *recombine* to form a water molecule.

$$H^+ + [OH]^- \rightarrow H_2O$$

To summarize what happens in pure water:

$$H_2O \rightleftarrows H^+ + [OH]^-$$

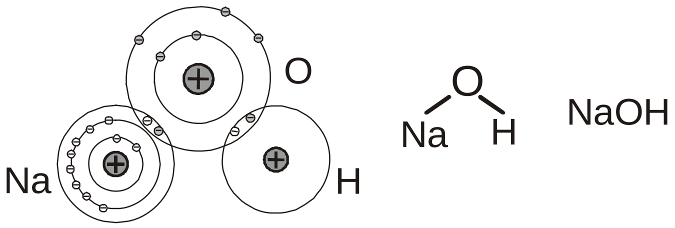
The double arrow means that both the forward and reverse reactions are occurring simultaneously. The emphasis of this reaction lies on the left hand side. Only one in half a billion water molecules undergoes dissociation.

As you can read from the chart, sodium has a very small electronegativity of 0.9 (compared to hydrogen). So the hydrogen ion would take over this single electron of the sodium atom and become a hydrogen atom. Two hydrogen atoms join up to built hydrogen gas.

$Na + H^+ \rightarrow Na^+ + H$

$2Na + 2H^+ \rightarrow 2Na^+ + H_2$

In a water molecule hydrogen can be substituted by a sodium atom. We get sodium hydroxide (NaOH).



In water sodium hydroxide (NaOH) will *dissociate* into Na^+ and $[OH]^-$.

$NaOH(s) \rightarrow Na^+ + [OH]^-$

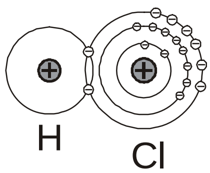
Na^+ is the **sodium ion**

$[OH]^-$ is called **hydroxide ion**

NaOH is called **sodium hydroxide**

As we will see later on, the hydroxide ion is an essential constituent of bases.

Hydrogen and chlorine can undergo covalent bonding, forming ***hydrogen chloride*.**



As chlorine has a high electronegativity level, the molecule will be highly polarized.

Dissolved in water, hydrogen chloride will dissociate completely forming ***hydrochloric acid.***

$HCl(g) \rightarrow H^+ + Cl^-$ (hydrochloric acid)

Hydrochloric acid is a strong acid.

Questions

1. Water molecules are polarized. What does this statement mean? Why are water molecules polarized?

2. Why is water an excellent solvent?

3. Does sugar dissociate in water?

.

4. Sodium hydroxide dissociates in water. What does this statement mean?

5. What is meant by dissociation of water?

Key Points

- Electronegativity is a quantitative measure of an atom's tendency to attract electrons in a chemical bond.
- Water molecules dissociate into *hydrogen ions* (H^+) and *hydroxide ions* ($[OH]^-$) ions.
- H^+ is an essential constituent of **acids.**
- $[OH]^-$ is a essential constituent of **bases.**
- HCl is a gas and is called *hydrogen chloride.*
- HCl dissociated in water forms *hydrochloric acid.*

8 Acids and Bases

In the last chapter I've already mentioned that water molecules dissociate into *hydrogen ions* (H^+) and *hydronium ions* ($[OH]^-$). But only to a very small extent.

$$H_2O \rightleftarrows H^+ + [OH]^-$$

There is only one hydrogen ion and one hydronium ion within half a billion water molecules.

We can increase the amount of hydrogen ions in water by adding hydronium chloride (HCl), resulting in hydrochloric acid.

If the concentration of hydrogen ions *increases*, the concentration of hydronium ions would *decrease*.

Why? Because the higher the concentration of hydrogen ions the greater the likelihood that a hydronium ion will find a hydrogen ion to recombine into a molecule of water.

$$H^+ + [OH]^- \rightarrow H_2O$$

Vinegar has a higher concentration of hydrogen ions than pure water has. This is also true of fruit juices, especially of lemon juice. It's also true of rain water.

If in a liquid the concentration of hydrogen ions (H^+) is <u>greater</u> than the concentration of the hydronium ions ($[OH]^-$), the liquid is called an **acid.**

Acids taste sour (but you should only taste weak acids such as vinegar or lemon juice).

On the other hand, you can also increase the amount of hydrogen ions by adding sodium hydroxide (NaOH). This would cause the concentration of hydrogen ions to decrease.

If in a liquid the concentration of hydrogen ions (H^+) is <u>less</u> than the concentration of the hydronium ions ($[OH]^-$), the liquid is called a **base.**

Bases feel soapy and indeed, soaps are basic.

To summarise the important points:

- **Acids** have a **high concentration of hydroxide ions (H^+)** and a low concentration of hydronium ions ($[OH]^-$). Acids taste **sour.**
- **Bases** have a low concentration of hydroxide ions (H^+) and a **high concentration of hydronium ions ($[OH]^-$).** Bases feel **soapy.**
- If a substance is neither *acidic* nor *basic*, it is considered **neutral.** Pure water is neutral.

The **pH scale**

The **pH scale** is used to indicate the **level of acidity or basicity** of a liquid.

At the moment it would be too difficult to explain how the pH scale is defined.

You only have to learn that:

- The pH scale goes from 0 to 14.
- The pH of pure water is 7.
- The pH of acids is less than 7.
- The pH of bases is greater than 7.
- The pH can be measured by indicators.

NaOH	14	strongly basic
	13	
	12	mildly basic
Washing soda	11	
	10	
Toothpaste	9	
Seawater	8	
Distilled water	**7**	**neutral**
Rain water	6	mildly acidic
Human skin	5	
Lemon juice	4	
Vinegar	3	
Battery acid	2	strongly acidic
HCl in stomach	1	
	0	

Indicators

Indicators are substances that change colour depending on whether they are in acidic or basic solutions.

Litmus indicator

- Turns blue in a base
- Turns red in an acid

Universal indicators change to a variety of colours depending on the pH level of the substance being tested.

The most important acids in chemistry are:

HCl	Hydrochloric acid
H_2CO_3	Carbonic acid
H_2SO_4	Sulfuric acid
HNO_3	Nitric acid

Hydrochloric acids dissociate into *hydrogen ions* (H^+) and *chloride ions* (Cl^-).

$HCl \rightarrow H^+ + Cl^-$

Sulfuric acids dissociate into *hydrogen ions* (H^+) and *sulphate ions* ($[SO_4]^{2-}$).

$H_2SO_4 \rightarrow 2H^+ + [SO_4]^{2-}$

Nitric acid dissociates into *hydrogen ions* (H^+) and *nitrate ions* ($[NO_3]^-$).

$HNO_3 \rightarrow H^+ + [NO_3]^-$

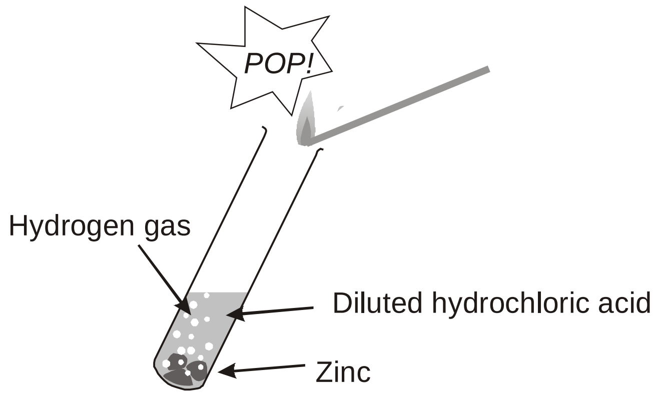
The pH of those acids depends on how heavily they are diluted. When concentrated they are strong acids.

Reactions of acids and metals

Experiment:

Metallic zinc in diluted *hydrochloric acid* produces *zinc ions* and *hydrogen* gas.

$2H^+(aq) + Zn(s) \rightarrow Zn^{2-}(aq) + H_2(g)$



The electronegativity of hydrogen (H) is 2.2, the electronegativity of zinc (Zn) is 1.6. So H^+ is able to “rob” electrons from the zinc. The solid zinc is transformed into zinc ions (Zn^{2-}), the hydrogen ion (H+) is transformed into atomic hydrogen. Every two hydrogen atoms combine into a hydrogen molecule (H_2).

Proof of hydrogen gas: Hydrogen gas can be lit by a flaming splint.

Using **gold instead of zinc**, no hydrogen gas would be produced. Gold has an electronegativity of 2.4 which is higher than that of hydrogen.

Acids “attack” metals, acids are **corrosive.**

Neutralisation

When an acid and a base react in an accurate way, salt and water are produced. Both are neutral, pH = 7.

Example

Hydrochloride acid (HCl) reacting with *sodium hydroxide* (NaOH) will produce *sodium chloride* (NaCl) and water (H_2O).

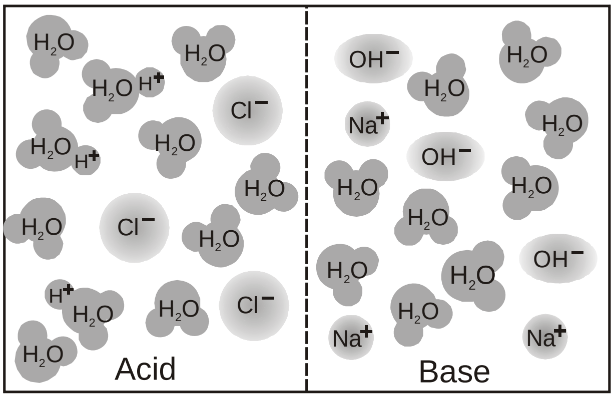
$$HCl + NaOH \rightarrow NaCl + H_2O$$

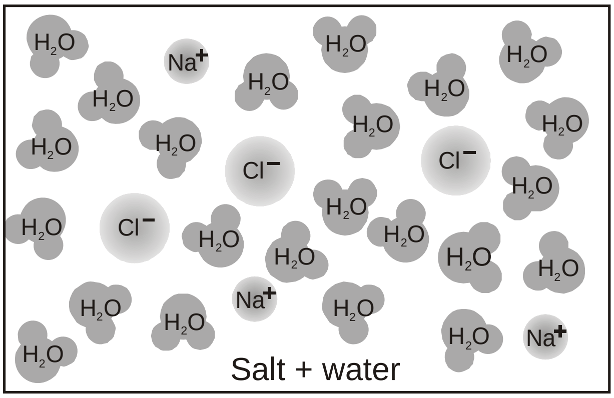
The same equation in ionic form:

$$H^+ + Cl^- + Na^+ + OH^- \rightarrow Na^+ + Cl^- + H_2O$$

The hydroxide ions (H^+) and the hydronium ions (OH^-) will join together to form water.

$$H^+ + OH^- \rightarrow H_2O$$





Example 1

Hydrochloride acid (HCl) plus *calcium hydroxide* produces *calcium chloride* ($CaCl_2$) plus water.

$$2HCl + Ca[OH]_2 \rightarrow CaCl_2 + 2H_2O$$

$$2H^+ + 2Cl^- + Ca^{2+} + 2[OH]^{2-} \rightarrow Ca^{2+} + 2Cl^- + 2H_2O$$

NaCl and $CaCl_2$ are **salts of the *hydrochloride acid.***
A salt of the hydrochloride acid is called a **chloride.**

NaCl is called **sodium chloride**

$CaCl_2$ is called **calcium chloride**

Example 2

Carbonic acid (H_2CO_3) and *sodium hydroxide* (NaOH) react to form ***sodium carbonate*** (Na_2CO_3) plus *water* (H_2O).

$$H_2CO_3 + 2NaOH \rightarrow Na_2CO_3 + 2H_2O$$

$$2H^+ + CO_3^{2-} + 2Na^+ + 2OH^- \rightarrow 2Na^+ + CO_3^{2-} + 2H_2O$$

A salt of the carbonic acid is called **carbonate.**

Example 3

Sulfuric acid (H_2SO_4) and *sodium hydroxide* (NaOH) react to ***sodium sulfate*** (Na_2SO_4) plus *water* (H_2O).

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

$$2H^+ + SO_4^{2-} + 2Na^+ + 2OH^- \rightarrow 2Na^+ + SO_4^{2-} + 2H_2O$$

A salt of the sulfuric acid is called **sulfate.**

Example 4

Nitric acid (HNO_3) and *sodium hydroxide* (NaOH) react to form ***sodium nitrate*** ($NaNO_3$) plus *water* (H_2O).

$$HNO_3 + NaOH \rightarrow NaNO_3 + H_2O$$

A salt of the nitric acid is called **nitrate.**

An acid plus a carbonate

an acid + a carbonate $\rightarrow$ salt + water + carbon dioxide gas

Hydrochloride acid (HCl) plus calcium carbonate ($CaCO_3$) produces *calcium chloride* (CaCl2) plus *water* plus *carbon dioxide* gas.

$$2HCl + CaCO_3 \rightarrow CaCl_2 + H_2O + CO_2(g)$$

$$2H^+ + 2Cl^- + Ca^{2+} + CO_3^{2-} \rightarrow Ca^{2+} + 2Cl^- + H_2O + CO_2$$

$$2H^+ + CO_3^{2-} \rightarrow H_2O + CO_2\ (gas)$$

The names of the acids and their salts

HCl	hydrochloric **acid**
NaCl	sodium **chloride**
Ca_2Cl	calcium **chloride**
H_2CO_3	**carbonic acid**
Na_2CO_3	sodium **carbonate**
$CaCO_3$	calcium **carbonate**
H_2SO_4	**sulfuric acid**
Na_2SO_4	sodium **sulfate**
$CaSO_4$	calcium **sulfate**
HNO_3	**nitric acid**
$NaNO_4$	sodium **nitrate**
$Ca(NO_4)_2$	calcium **nitrate**

H	hydrogen	**H^+**	hydrogen ion
OH^-	hydroxide ion		

Questions

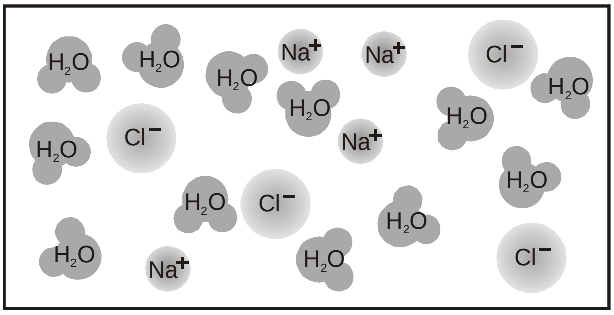
1. List the properties of acids.

2. What is an indicator?

3. What is a neutralisation reaction? Give one example.

4. Give two examples of household acids and household bases.

5. What does the picture represent?



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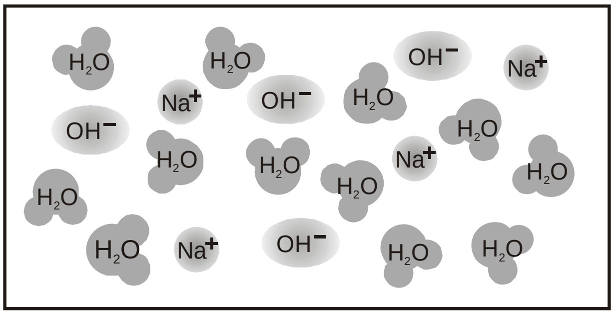
6. What does the picture represent?



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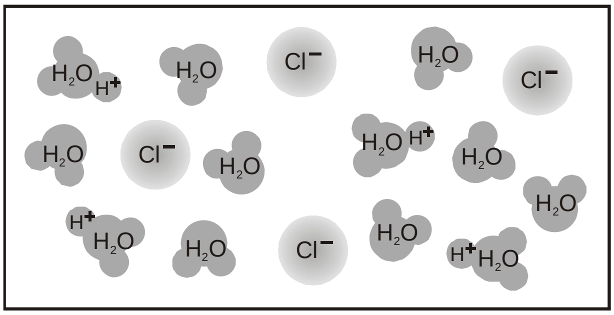
7. What does the picture represent?



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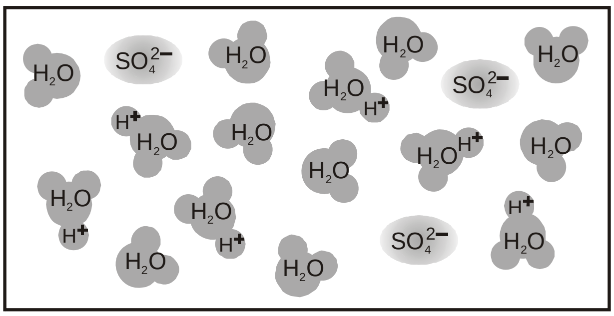
8. What does the picture represent?



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9. A salt of the hydrochloric acid is called ………………

   A salt of the carbonic acid is called ………………..……

   A salt of the sulfuric acid is called ……………………

   A salt of the nitric acid is called ………………….……

10. NaCl is called ……………………………..……………

    $CaCO_3$ is called ……………………..…………………

    K_2SO_4 is called ……………………………..…………

    $AgNO_3$ is called ………………………..………………

Key Points
• The pH scale runs from 0 to 14.
• An acid is a substance which has a pH < 7, turns blue litmus red and has a sour taste.
• A base is a substance which has a pH > 7 and which turns red litmus blue.
• An indicator is a chemical which shows, by means of colour change, whether a substance is acidic, basic or neutral.
• A neutralisation reaction occurs when an acid reacts with a base, forming salt and water.
• acid + metal $\rightarrow$ salt + water + H_2(g)
• acid + carbonate $\rightarrow$ salt + water + CO_2(g)