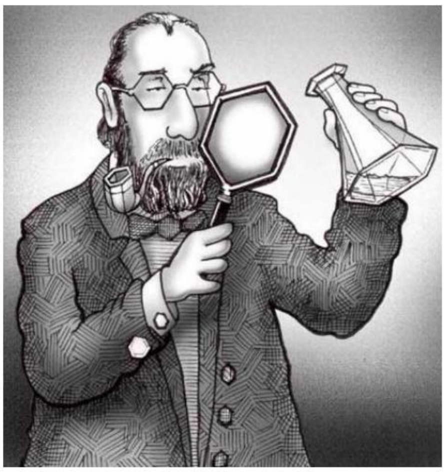
Sint-Barbaracollege, Gent
Department of Sciences

VADEMECUM CHEMISTRY 2ND DEGREE



@chemistry department

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1 PURE SUBSTANCES AND MIXTURES

1.1 Matter and pure substances

Objects are made up of **substances.** We call all substances together **matter.** Chemistry is the study of substances and their specific properties or substance constants.

A **pure substance** consists of one type of dust particles. A **mixture** is always made up of several pure substances.

We can recognize a pure substance by its physical constants such as aggregation state at room temperature, density, boiling point, melting point, solubility,
etc. Pure substances occur in nature but can also be made artificially (synthetically). However, the properties are the same regardless of the preparation method.

1.2 Types of mixtures

A mixture is therefore a collection of several substances mixed together. Mixtures in which different types of particles can be distinguished 'by sight'[1] are called **heterogeneous mixtures .** In **homogeneous mixtures,** the substance properties are the same in all places.

Colloidal mixtures are a transitional form between homogeneous and heterogeneous mixtures.

There are many types of mixtures. The most important mixtures used in chemistry are the following:

Heterogeneous mixtures:

suspension: This is a mixture in which a solid does not dissolve in a solvent but is distributed as fine grains through the solvent (water and sand). A suspension is cloudy and over time the substances will separate based on their difference in density.

emulsion: This is a mixture of two immiscible liquids (water and oil). An emulsion separates itself after a short or longer time. You then get two layers, with the liquid with the highest density forming the bottom layer.

Other heterogeneous mixtures are: foam (gas finely divided in a liquid), mist (liquid that is finely divided in a gas) and smoke (finely divided solid in a gas)

Homogeneous mixtures:

Solution: This is a mixture of a liquid (the solvent) and another substance (or substances). In a solution, the substances are completely miscible with each other and you therefore get a homogeneous mixture.

solid – liquid: sugar in water
liquid – liquid: alcohol in water
gas – liquid: oxygen gas in water

Characteristic of a solution is that it is always clear. You can see through it. Solutions are therefore transparent, which is not the same as colorless!

Other homogeneous mixtures are: alloys (mixture of metals).

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1 The boundary between homogeneous and heterogeneous mixtures is not always easy to define and often depends on the optical aids used (microscope or magnifying glass).

1.3 Separation of mixtures

Mixtures can be separated back into their original components.
Separation techniques rely on specific physically different properties of the pure substances that make up the mixture (boiling point, density, solubility, etc.).
It is important to realize that no new substances are created during separation. Separation is therefore not a chemical reaction (see 6).

technology	aggregate state	examples	
more right	vl – vl	oil and water	
decantation	v–vl	sediment in wine	
filtration	v–vl	sand and water	
centrifugation	v – vl vl – vl	dry spinning lettuce skimming milk caffeine	
extraction	v – v v–vl	from coffee oil from peanuts	
adsorption	v–vl g – vl g – g	coloring from wine, fragrance from drinking water, gas	
distillation	vl – vl	mask, alcohol and water	
crystallization	v–vl	extraction of sea salt	
chromatography	v – v g – g vl – vl	dyes in felt-tip pen purification uranium organic synthesis	

In bold are the separation techniques that were certainly seen in the third year.

2 THE CONSTRUCTION OF THE FABRIC

2.1 Classification of substances

In chemistry, **pure substances** are studied. A pure substance is made up of one type of particles. These can be either molecules or atoms.

A **compound substance** can be further decomposed into two or more other substances.
A **single substance** can no longer be further decomposed into other substances.

2.2 Molecules and atoms (elements)

Molecules are made up of several **atoms.** We also call the different types of atoms the **elements.**

All simple substances are made up of only one type of atom (iron: Fe, dioxygen: O_2, etc.).

Composite substances are made up of different types of atoms (water: H_2O, ammonia: NH_3, carbon dioxide: CO_2, etc.).

There are three *types* of elements: **metals** (iron: Fe, lead: Pb, mercury: Hg…), **non-metals** (carbon: C, hydrogen: H, sulphur: S…) and **noble gases** (helium: He, argon: Ar…).

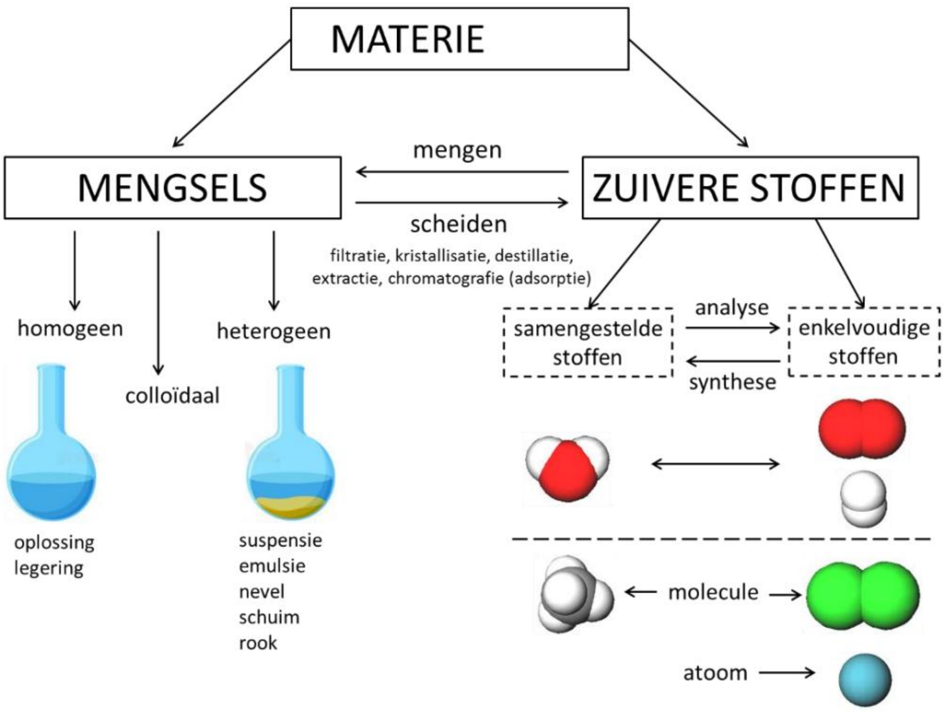
Each element is represented by a specific **symbol.**
The notation 'Cu' has two meanings:

- the atomic type or element copper
- 1 atom of copper

The elements to be known after the second degree are shown below:

METAL		into a career	IN
lithium	That	plutonium	Could
sodium	Already	polonium	After
potassium	K	NON-METALS	
beryllium	Be	hydrogen	H
magnesium	Mg	boron	B
calcium	That	carbon	C
barium	Not	nitrogen	N
aluminium tin	Al	oxygen	O
	Sn		F
antimoon lood	Sb	fluorine	And
	Pb	silicon	P
chroom	Cr		S
manganese	Mn	phosphorus sulfur chlorine	Cl
iron	Fe	arsenic	As
cobalt	Co	selenium	Se
nickel	In	bromine	Br
copper	With	iod	I
zinc	Zn	NOBLE GASES	
silver	At	helium	He
gold	At	neon	Yes
platinum	Pt	argon	With
cadmium	Cd	krypton	NOK
	Hg	xenon	Car
mercury germanium	Ge	radon	Rn

Overview diagram:



2.3 Atomic construction

2.3.1 Components of an atom

An atom is, roughly speaking, made up of a central, solid **nucleus** with a **positive charge,** surrounded by an 'extensive' thin space **(mantle)** in which the **negatively charged electrons** are located. Research showed that the nucleus **is** made up of 2 types of nuclear particles **(nuclides): protons** and **neutrons.**

particles	*symbol*	*place*	*charge (relative) mass*	*(relative)*
of proton	p^{+}	core	+ 1	1
neutron	n^{0}	kern	0	1
electron	It is	mantel	- 1	1/1840

2.3.2 Atomic number and mass number

An atom is electrically neutral: the number of positive charges is equal to the number of negative charges. In other words: the number of protons in the nucleus is always equal to the number of electrons in the electron shell.

The **atomic number Z** indicates the number of protons; it is characteristic of the element.
The **mass number A** is equal to the sum of the number of protons and neutrons.
A – Z equals the number of neutrons.

notation:

massage number ← A
X → symbol element
atomic number ← WITH

Since Z is characteristic of the element, it can also be omitted from the notation.

example: 23Na has Z = 11

number of p+ = 11
number of n0 = 12
number of e- = 11

number of nuclides = 23

2.3.3 The electron configuration

In the electron shell, the electrons are divided over 7 **shells** or **energy levels** (atomic model of Bohr).

The peels are indicated by a letter: KLMNOP Q.

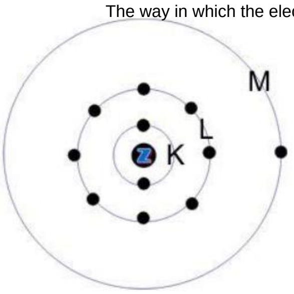
They are characterized by a **main energy level,** which allows you to calculate the maximum number of electrons per shell (= 2n2).

shells main	K	L	M	N
energy level = n maximum number of electrons = 2n2	1 2	2 8	3 18	4 32

The electrons of an atom are always at the lowest possible shell or energy level.

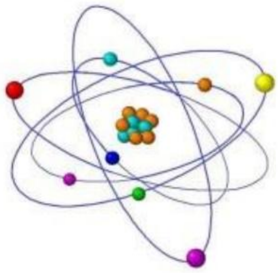
However, the number of electrons in the outer shell can never exceed 8!

The way in which the electrons are distributed among the different shells is called the **electron configuration .**



The electrons found in the outer shell of an atom **are called valence electrons .**

Always keep in mind that an atom occupies a 3-dimensional space and is not a flat object. The figure on the left is therefore a flat projection of reality[2]



The electron configurations of the first 18 elements of the periodic table can be found in the table below.

Atomic number, Z Element 1 2 3 4 5 6 7 8 9 10 11 12		K	L	M	N
	H		-	-	-
	He	1	-	-	-
	That	2		-	-
	Be	2	1	-	-
	B	2	2	-	-
	C	2	3	-	-
	N	2	4	-	-
	O	2	5	-	-
	F	2	6	-	-
13	Yes	2	7	-	-
14	Already	2	8		-
15	Mg	2	8	1	-
16	Al	2	8	2	-
17	And	2	8	3	-
18	P	2	8	4	-
	S	2	8	5	-
	Cl	2	8	6	-
	With	2 2	8 8	7 8	-

2 Bohr's atomic model is a useful but also incomplete atomic model. This will be further refined in the fifth year.

3 THE PERIODIC SYSTEM (PSE)

In the periodic table, the elements are arranged according to their increasing atomic number (from left to right).
Horizontally, the periodic table is divided into seven **periods.** The **groups** are listed vertically.

The elements of **the same group** always exhibit **similar chemical properties** because they have the same number of valence electrons (electrons on the outer shell).

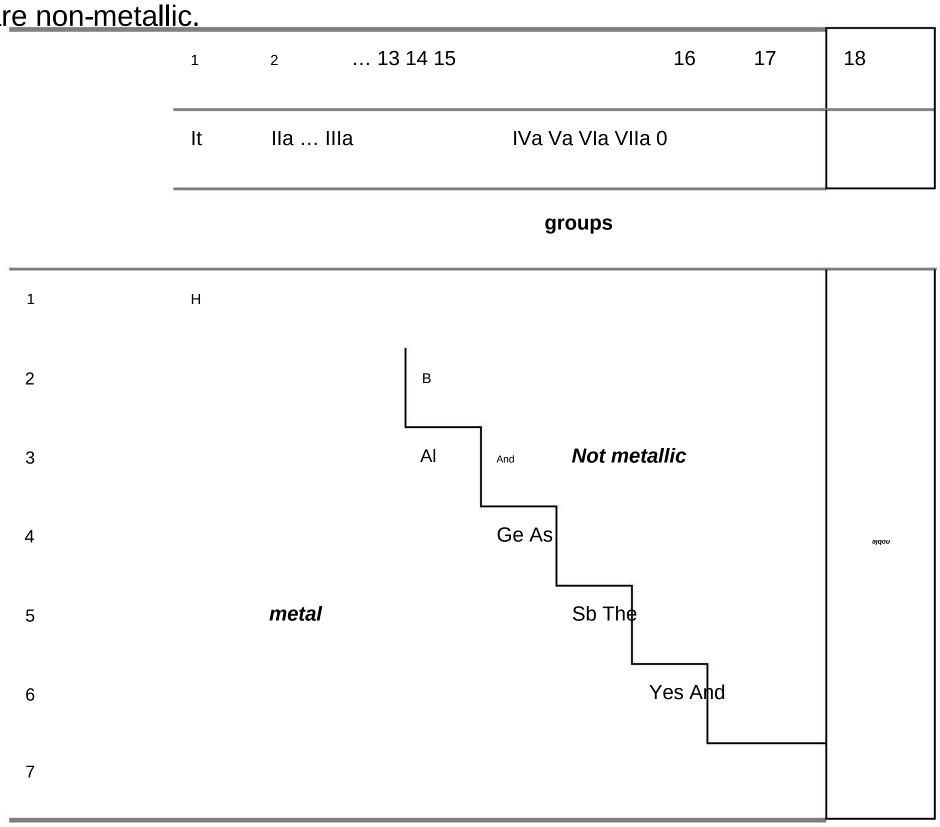
The **period** represents the **number of filled shells** (or energy levels) of the element.

In recent versions of the PSE, the groups are numbered consecutively from 1 to 18. In older versions, Roman numerals with letters are used to distinguish main and minor groups (a and b groups). The Roman numeral represents the number of valence electrons. The group '0' is the special group in the periodic table.
This contains the noble gases. These have 8 valence electrons and are extremely stable.

The main groups (a-groups) are given a name:

Ia 1 alkali metals
IIa 2 the alkaline earth metal
IIIa 13 the aard metal
IVa 14 carbon group
Va 15 nitrogen group
VIa 16 oxygen group
VIIa 17 halogen
0 18 noble gases

Most elements have a metallic character, only hydrogen and the elements in the upper right corner of the table are non-metallic.



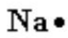
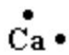
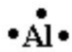
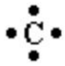
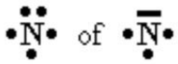
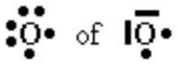
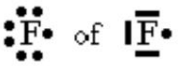
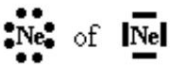
4 CHEMICAL BONDS

4.1 The octet structure

When forming chemical bonds, each atom will strive to get eight electrons (= **octet structure** or **noble gas configuration)** in its outer shell. An exception to this is the element H, which emulates the noble gas configuration of He (only 2 valence electrons).

In chemical processes, only the outer electrons (the **valence electrons)** play a role.

With **Lewis notation ,** the valence electrons are depicted as dots around the symbol. From five electrons onwards, dashes are also used to represent electron pairs.

element	*valence electrons*	**4.1.1 Lewisstruct o'clock**
Already	1	
That	2	
Al	3	
C	4	
N	5	
O	6	
F	7	
Yes	8	

4.2 The concept of 'ion'

Ions are charged particles (made up of one or more atoms) and are created when one or more electrons are added or removed from an atom.

There are two types of ions: positive ions **(cations,** electrons taken away) and negative ions **(anions,** electrons taken away). The formation of an ion from an atom depends on the number of valence electrons.

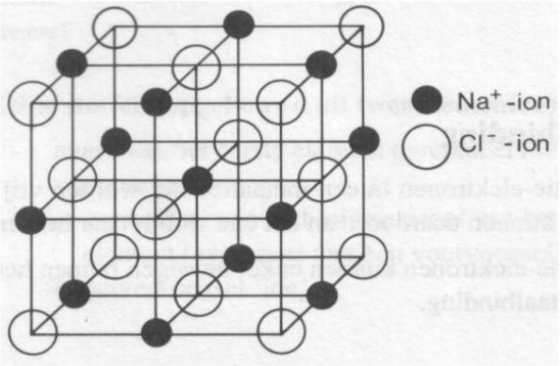
element	valence electrons	number of electrons given up or taken up	ion	name of the ion
Already	1	1 surrendered	Na^+	natriumion
Mg	2	2 surrendered	Mg^{2+}	magnesiumion
Al	3	3 conceded	Al^{3+}	aluminiumion
O	6	2 included	O^{2-}	oxide-ion
S	6	2 included	S^{2-}	sulfide-ion
F	7	1 included	F^-	fluoride-ion

4.3 De ionbinding

Metal atoms lose electrons very easily. This creates positive ions.
Non-metal atoms gain electrons. This creates negative ions.
These positive and negative ions exert an electrical attraction on each other.
This creates an ionic bond.

In an ionic bond, the ions will arrange themselves in a regular pattern.
We call this arrangement of ions an **ionic lattice.**

The chemical formula NaCl does not actually represent a molecule in itself, but is the unit that repeatedly returns in the lattice of this substance.



4.4 The covalent bond or atomic bond

Nonmetals can achieve the noble gas configuration by sharing electrons. A shared electron pair forms a covalent bond.
The covalent bond can be easily represented with the Lewis structure.

example: formation of dichlor:

remark:
A rule is often used (difference in electronegative values (see 5.3) of two neighboring atoms > or < 1.6) to indicate the difference between an ionic bond and a covalent bond. However, this rule does not always apply and it is therefore better to use the rule of thumb that an ionic bond is formed between metals and non-metals and a covalent bond is formed between non-metals themselves.

4.5 The metallic bond

The valence electrons in a metal lattice are attracted quite weakly to the nucleus.
They can therefore move quite easily between the atoms and are also called 'free' electrons.
The charge distributions that result keep the metal atoms bound.

5 MOLECULAR FORMULAS AND NAMES

5.1 Molecuulformules

Only pure substances can be represented by a chemical formula.
In the formula, the index indicates the number of atoms of a particular element.
A coefficient indicates the number of molecules.

3 H O2

coefficient index

In this example the following quantities are present:
3 molecules of water, 6 atoms of H and 3 atoms of O

A formula is characteristic of a pure substance but independent of the state of aggregation.

Naming:

Systematic (or scientific) names are internationally agreed upon; trivial or usage) names usually come from industry. For the naming of chemical formulas, reference is made to 7 and 8.

Prefixes are used according to the index in the molecular formula.

index 1	*prefix*
	mono
2	Of
3	three
4	tetra
5	penta
6	hexa
7	hepta
8	octa
9	for him
10	deca

5.2 Electronegative value – oxidation number (see also 11)

5.2.1 Electronegative value (EN)

The electronegative value is a measure of the force with which an atom attracts electrons.

Linus Pauling has drawn up a scale from 0.7 to 4.0. These electronegative values are listed in the periodic table. Metals have a low EN value and are more likely to reject electrons to achieve the octet structure. Non-metals have a high EN value and will more easily attract additional electrons.

5.2.2 Oxidation number (OG)

Each element is characterized by a certain oxidation number in a specific compound.

The oxidation number is represented by a Roman numeral preceded by a + or a – sign.

5.2.3 Oxidation number for ionic bonds

Metal atoms are 'electropositive' because they very easily form positive ions by releasing electrons. The positive charge of that ion is the oxidation number of the element involved.

Nonmetal atoms are 'electronegative' because they form negative ions by absorbing electrons. The negative charge of the ion is called the oxidation number of the element.

element	*ion*	*oxidation number*
Already	Na+	+ I
Mg	Mg2+	+ II
O	O2-	- II
F	F -	- I

5.2.4 Oxidation number for covalent bonds

We find the oxidation number of the elements involved by completely shifting the bond electrons towards the element with the largest electronegative value. We then look at which partial charge (or partial charge) has arisen on the atoms.

example: CO2

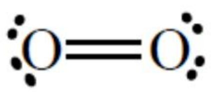
Electronegative values are: 3.5 2.5 3.5

Part loads are: ÿ 2- ÿ 4+ ÿ 2-

The oxidation numbers are: -II +IV -II

If there is no difference in the electronegative value, then there is no charge shift and each atom is assigned half of the bonding electrons.

example: O2



Electronegative values	3,5	3,5
The oxidation numbers are:	0	0

5.2.5 Oxidation number for metallic bonds

Metal atoms in a metal lattice have an oxidation number equal to zero.

Some rules apply to determining the oxidation numbers of elements:

- for the atoms of a simple substance the OG = 0.
- for a neutral compound the sum of the OG = 0.
- for monatomic ions the OG = the relative ionic charge.
- for polyatomic ions, the sum of the OG = the relative ionic charge.
- the OG of an oxygen atom in a compound is usually -II.
- the OG of a hydrogen atom in a compound is usually +I.

For elements from the side groups and in case of multiple values for elements from the main groups, a limited table with OG is consulted to write formulas (see 7)

6 CHEMICAL REACTIONS

6.1 Difference between physical and chemical phenomena

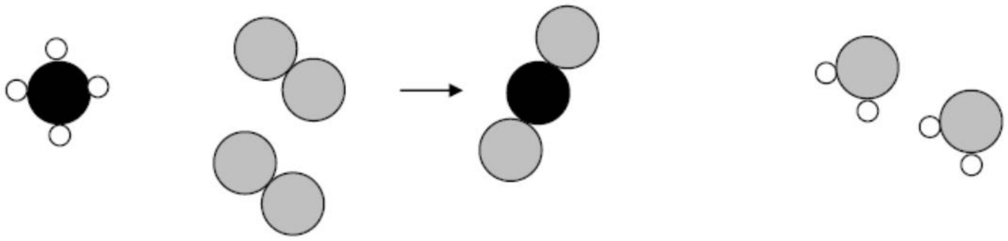
During a physical phenomenon, no new substances are formed (e.g. evaporation of water, separation techniques). With a chemical phenomenon, new substances are formed (e.g. combustion of wood).

6.2 Chemical reaction equations

Reaction is essentially the arranging and/or rearrangement of atoms to form new molecules. Chemical reactions arise because atoms want to achieve the octet structure (8 valence electrons).
The reactions are noted in reaction equations: reactants ÿ reaction products

In the reaction equation, the number of atoms of a certain species must be equal before and after the reaction. We can correct this by using the correct **coefficients** (smallest whole numbers).

1 CH_4 (g) + **2** O_2 (g) ÿ **1** CO_2 (g) + **2** H_2O (vl)



The coefficient 1 is usually omitted: CH_4 (g) + 2 O_2 (g) ÿ CO_2 (g) + 2 H_2O (vl)
The letters in brackets after each formula indicate the state of the substance in the reaction (g = gas, v = solid, vl = liquid and sol = in solution).

To draw up a chemical reaction equation, we always follow the steps below (see also 7.5):

1 Recognize reaction pattern
2 Write correct chemical formulas (use indices)
3 Adjust the reaction equation (use coefficients)

6.3 Laws that determine the course of a chemical reaction

Lavoisier's law: During a reaction in a closed space, the total mass of the substances involved remains unchanged.

Proust's law: When two or more substances react with each other, this always happens in a constant mass ratio.

As you can see from the figure above on this page, no atoms disappear during a chemical reaction!

6.4 Types of reactions

6.4.1 Analysis or decomposition reaction

compound substance ÿ two or more substances		
A	ÿ	**B + C**

thermolysis: decomposition by heat

photolysis: decomposition by light

electrolysis: decomposition by electric current

6.4.2 Synthesis or compound reaction

fabric + fabric ÿ composite fabric
A + B ÿ C

6.4.3 Substitution

An element in a substance is replaced by another element.

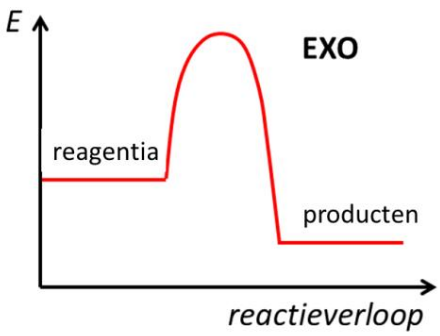
AB + C ÿ AC + B

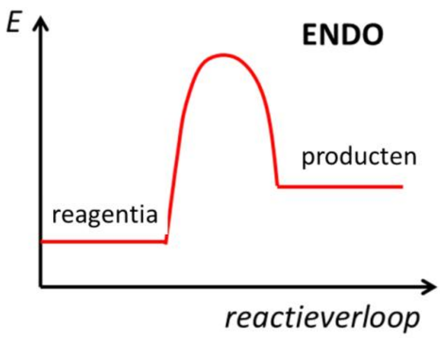
6.5 Energy aspect

An **exothermic** reaction is a chemical reaction in which energy is released in the form of heat.

An **endothermic** reaction is a chemical reaction in which heat is extracted from the environment.

These reactions can be represented in an energy diagram. This shows that in an exothermic reaction the products formed have a lower energy content than the original reactants. In an endothermic reaction, the reactants absorb energy and the products formed will have a higher energy content.





remark:
The energy released or absorbed can take a different form of energy than heat, such as light. In this case we speak of an exo-energetic or endo-energetic reaction.

7 INORGANIC MATERIALS[3]

7.1 They oxidize

Metal oxides are created by the combustion of metals. They have the general formula **MxOy** and are solids at room temperature.

Non-metal oxides are formed by the combustion of non-metals. They have the general formula **nMxOy** and are gaseous at room temperature.

When naming we first name the metal or non-metal and then the ending 'oxide' (possibly with prefixes).

examples:	CaO calciumoxide	K2O dipotassiumoxide
	NO nitric oxide	P2O5 diphosphorpentaoxide

7.2 Hydroxiden

A hydroxide is formed during the reaction between a soluble metal oxide and water. The general formula is **M(OH)x.**

The formula of a hydroxide consists of two parts:

metal ← hydroxide → one or more hydroxide groups

When naming we first mention the metal and then the ending 'hydroxide' (possibly with a prefix).

examples:	KOH	potassium hydroxide Ca(OH)2 iron	calciumdihydroxide
	Fe(OH)2	dihydroxide Fe(OH)3	iron trihydroxide

Hydroxides exhibit the following characteristics:
- ÿ they have a soapy taste
- ÿ only some are easily soluble in water (see 9)
- ÿ the soluble hydroxides are good conductors of electric current in aqueous solutions
- ÿ the basic properties are due to the presence of OH- ions in an aqueous solution
- ÿ hydroxides are ionic bonds and are therefore solid at room temperature

When we dissolve a hydroxide in water, it splits into ions: **dissociation.**
example: Ca(OH)2 (v) ÿ Ca2+ (sol) + 2 OH- (sol)

An ammonia solution also has basic properties. It yields OH ions according to the reaction:

$$NH_3 + H_2O \text{ ÿ } NH_4^+ + OH^-$$

The NH_4^+-ion is called the ammonium ion (NH4OH = ammonium hydroxide).

---

3 M = metal, nM = non-metal

7.3 Acids

An acid is formed from the reaction between a non-metal oxide and water. The general formula for acids is **HxZ.**

The formula of an acid consists of two parts:

pickles

one or more hydrogens acid residue

Acid residues are always negative ions.

There are two types of acids.

Binary acids consist of H and a non-metal: **HxnM** (HCl, H_2S, ...)

Ternary acids always consist of H, a non-metal and O: **HxnMOy** (HNO_3, H_2SO_4, ...)

Some important acids with their acid residues

pickles	name of the acid	trivial name acid resid	ue name of the	acid residue
HF	**waterstoffluoride**		**F^-**	**fluoride-ion**
HCl	**waterstofchloride**	**hydrochloric acid**	**Cl^-**	**chloride-ion**
HBr	**waterstofbromide**		**Br^-**	**bromide-ion**
HI	**waterstofjodide**		**I^-**	**jodide-ion**
H_2S	**diwaterstofsulfide**		**S^{2-}**	**sulfide-ion**
HNO_3	**hydrogen nitrate**	**nitric acid**	**NO_3^-**	**nitrate ion**
H_2CO_3	**Dihydrogen carbonate**	**carbonic acid**	**CO_3^{2-}**	**carbonate ion**
H_2SO_4	**diwaterstofsulfate**	**sulphuric acid**	**SO_4^{2-}**	**sulfate ion**
H_3PO_4	**trihydrogen phosphate**	**phosphoric acid**	**PO_4^{3-}**	**phosphate ion**
$HClO_3$	**hydrogen chlorate**		**ClO_3^-**	**chloraat-ion**
HIO_3	**hydrogen iodate**		**IO_3^-**	**iodate ion to**
$HBrO_3$	**hydrogen bromate**		**BrO_3^-**	**bromate ion**

Acids exhibit the following characteristics: ÿ they are easily soluble in water ÿ in a very dilute aqueous solution they have a sour taste ÿ they react with many metals, releasing hydrogen gas ÿ they conduct electric current in an aqueous solution ÿ the acidic properties are due to the presence of H^+ ions in an aqueous solution ÿ acids are covalent bonds ÿ at room temperature binary acids are gaseous and ternary acids are liquid

When we dissolve an acid in water it forms ions: **ionization.** *examples:* HCl (g)ÿ H^+ (sol) + Cl^- (sol)

H_2SO_4 (vl) ÿ H^+ (opl) + HSO_4^- (opl) ÿ 2 H^+ (opl) + SO_4^{2-} (opl)

7.4 Salting

A salt is formed from a reaction between an acid and a base:

acid + base ÿ salt + water.

We call such a reaction a neutralization reaction.

The general formula for a salt is **MxZY.** A salt therefore consists of at least two parts:

salty

metal (or ammonium ion) acid residue

When naming we mention the name of the metal ion or ammonium ion, followed by the name of the acid residue (possibly with prefixes). Corresponding to the acids, binary **(MxnMy)** and ternary salts **(MxnMOy) also exist.**

examples:	KNO3 potassium nitrate	NaCl	sodium chloride
	MgSO4 magnesium sulfate K2S		dipotassiumsulfide
	Fe(NO3)3 iron trinitrate	Na2CO3	disodium carbonate
	(NH4)3PO4 tri-ammoniumfosfaat		

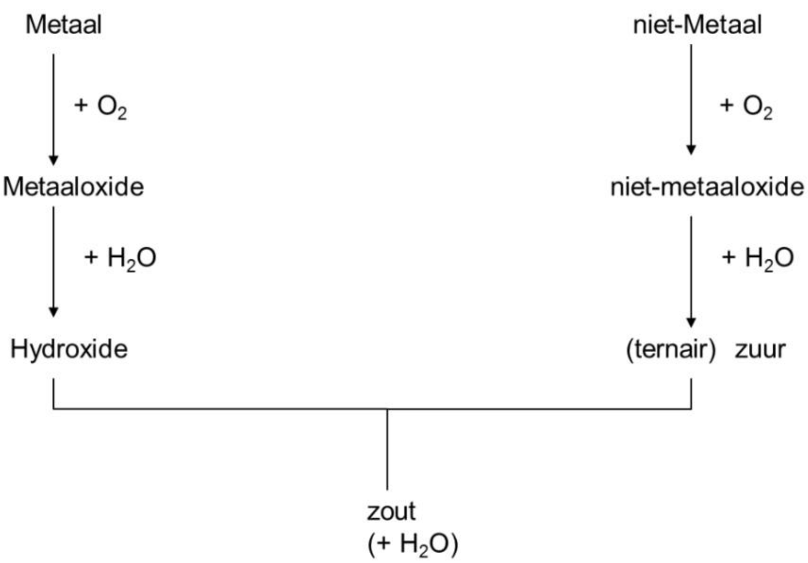
Salts exhibit the following characteristics:

- ÿ the solubility in water depends on the type of ions (see 11)
- ÿ When they are soluble in water, they conduct electric current well
- ÿ salts are ionic bonds and are therefore solids at room temperature

A salt splits into ions when dissolved in water: **dissociation**

example: Ca(NO3)2 (v) ÿ Ca2+ (sol) + 2 NO3 - (sol)

7.5 Relationship between the inorganic substances



Extra: metal + acid ÿ salt + H2

8 ORGANIC SUBSTANCES

8.1 Organic chemistry

ÿ All organic substances contain the element carbon.

ÿ The branch of chemistry concerned with the study of organic substances we call organic chemistry or carbon chemistry.

8.2 Has carbon atom

ÿ The element carbon has four electrons in its outer shell: it is therefore able to form four stable, covalent bonds.

ÿ Carbon has the unique property of forming chains, where the number of carbon atoms can be very high.

ÿ Carbon chains can be branched and can also contain multiple bonds. Moreover, there are also cyclic connections (subject matter 5th year).

8.3 Hydrocarbons (n-alkanes)

Structure: only C and H atoms, all atoms are singly bonded.

General gross formula: **CnH2n+2.**

Name: stem + 'on'.

No. C-at.	Structural formula	Gross formula	Name	aggregation -state
	CH4 methane gas	CH4		
1	CH3-CH3 ethane gas	C2H6		
2	CH3-CH2-CH3 propane gas	C3H8		
3	CH3-CH2-CH2-CH3 butane gas	C4H10		
4	CH3-CH2-CH2-CH2-CH3 liquid pentane	C5H12		
5	CH3-CH2-CH2-CH2-CH2-CH3 hexane liquid	C6H14		
6	CH3-CH2-CH2-CH2-CH2-CH2-CH3 heptane liquid	C7H16		
7	CH3-CH2-CH2-CH2-CH2-CH2-CH2-CH3 octane liquid	C8H18		
8	CH3-CH2-CH2-CH2-CH2-CH2-CH2-CH2-CH3 nonane liquid	C9H20		
9 10	CH3-CH2-CH2-CH2-CH2-CH2-CH2-CH2-CH2-CH3 C10H22	decane liquid		

remark:

Molecular formulas (see 5) of the substances in question can be displayed in two ways. The gross formula gives the number of atoms of each species and is the most concise form of writing. The structural formula provides information about the way in which the carbon atoms are bonded to each other.

As the number of carbon atoms in the chain increases, the boiling point of the alkanes also increases. Based on this principle, the different alkanes in petroleum can be separated from each other via fractional distillation.

8.4 Alcohols (alkanols)

Functional group: the OH group (hydroxyl group)

General formula: **CnH2n+1OH**

Name: stem + 'ol'; from propanol the position of the alcohol function is indicated by a number.

examples:

- CH3 – CH2 – OH ethanol
- CH3 – CH2 – CH2 – CH2 – OH butan-1-ol
- CH3 – CH2 – CH – CH3 butan-2-ol

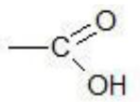
OH

Alcohols are easily soluble in water. In their pure form they are flammable:

CH3OH (vl) + O2 (g) ÿ CO2 (g) + H2O (vl)

8.5 Carboxylic acids (alkanoic acids)

Functional group: the COOH group (carboxyl group).



General formula: **Cn-1H2n-1COOH**

Name: stem + 'acid'; the carbon of the carboxyl group belongs to the main chain and is always terminal.

examples:

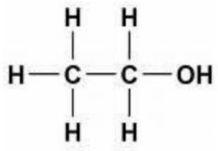
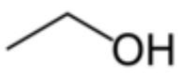
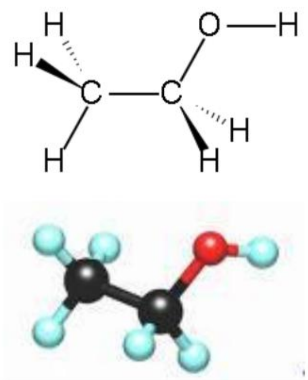
- HCOOH methanoic acid (formic acid)
- CH3-COOH ethanoic acid (acetic acid)

Carboxylic acids are organic acids and ionize in an aqueous medium:

CH3-COOH ÿ CH3-COO- + H+

remark:

There are different ways in which organic substances can be represented.
Below is an example for ethanol.

Gross formula	Structural formula	Sawtooth projection	3D representations
C2H6O			

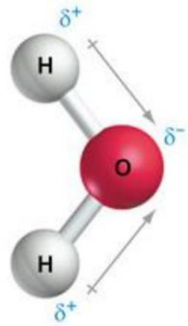
9 THE BEHAVIOR OF SUBSTANCES IN WATER

9.1 Solutions and solubility

A **solution** consists of a **solvent** and a **solute.**
A solvent is always a liquid. A solute can be liquid, solid or gaseous.

The solubility of a substance depends on:
ÿ the nature of the substance
ÿ the nature of the solvent.

Water as solvent



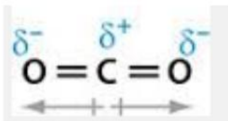
In a water molecule, the oxygen atom attracts the bonding electrons (higher EN value). It therefore becomes partially (partially) negatively charged while the hydrogen atoms become partially (partially) positively charged.

The water molecule is a polar molecule or **dipole molecule.**

Due to its polar character, the water molecule is sensitive to ions around it. The polar water molecules form a 'dipole shell' around ions, so that they remain in a dissolved state (denoted by (opl) in reaction equations).
A distinction must be made between the concepts of dissociation and ionization.
Dissociation occurs in ionic bonds (salts and hydroxides). The dipole forces of water pull the ions loose from their ionic lattice. **Ionization** occurs with polar, covalent bonds (acids)[4] . Ions are formed during this process.

As a general rule we can say that polar substances dissolve well in a polar solvent and non-polar substances dissolve well in a non-polar solvent.

remark:
A substance is polar when electron shifts occur between two elements due to a difference in EN value (e.g. HCl and NH_3). However, when the substance is constructed symmetrically, the electron shifts cancel each other out and the substance is still non-polar (e.g. CH_4 and CO_2)

Polar substance	Non-polar substance
$\overset{\delta^+}{H} \overset{\longrightarrow}{-} \overset{\delta^-}{Cl}$	 The two dipole vectors cancel each other out for symmetry

4 Metal oxides and non-metal oxides do not dissolve in water, but they do react with water to form new substances, namely hydroxides and acids.

9.2 Conductivity of solutions

Substances that, in a dissolved state (or molten state), completely or partially split into positive and negative ions, making the solution capable of conducting electric current, are called **electrolytes.** Substances that do not conduct electric current are called **non-electrolytes.**

Substances that completely split into ions in water are called **strong electrolytes .**
Substances that only partially dissociate are **weak electrolytes.**

Ionic compounds and some polar covalent compounds (acids) are electrolytes.
Non-polar covalent compounds are always non-electrolytes.

note: There
is no direct relationship between electrolyte strength and solubility of a substance in water. For example, AgCl and Ba(OH)2 are difficult to dissolve in water (see 10.2) but they behave as strong electrolytes. The small part of those substances that dissolves is completely dissociated. On the other hand, there are also many substances that dissolve well in water (sugar) but that do not exhibit electrolyte properties because no ions are formed.

9.3 Acidity

We can indicate the acidity of a solution by pH values. The pH values range from 0 to 14.

A pH value equal to 7 means that the solution is neutral; the number of H+ ions is equal to the number of OH- ions.

A pH value that is less than 7 indicates an acidic solution, where the solution contains more H+ ions than OH- ions.

If the pH value is above 7, we have a basic solution. Such a solution contains more OH- ions than H+ -

ions.

	effecten op het milieu	pH waarde	voorbeelden
ZUUR		pH = 0	batterijzuur
		pH = 1	zwavelzuur
		pH = 2	citroensap, azijn
		pH = 3	sinaasappelsap, soda
	alle vissen gaan dood (4,2)	pH = 4	zure regen (4,2-4,4) zuur meer (4,5)
	kikkereieren, kikkervisjes, rivierkreeften en eendagsvliegen gaan dood (5,5)	pH = 5	bananen (5,0-5,3) schone regen (5,6)
	regenboogforellen gaan dood (6,0)	pH = 6	schoon meer (6,5) melk (6,5-6,8)
NEUTRAAL		pH = 7	puur water
		pH = 8	zeewater, eieren
		pH = 9	zuiveringszout (zit in bakpoeder)
		pH = 10	magnesiummelk (zit in maagzuurremmers)
		pH = 11	ammonia
		pH = 12	zeepwater
		pH = 13	bleek
BASISCH		pH = 14	gootsteenontstopper

To qualitatively demonstrate the acidity of a solution, acid-base indicators can be used.

Indicator	Color red red colorless	**Cover area** 4.5 – 8	**Color at high pH**
Lakmoes	**at**	3.1 – 4.4	blue
Methyloranje	**low**	8.2 – 10	yellow
Phenolphthalein	**pH**		purple

To determine the pH more precisely (quantitatively), a mixture of indicators (e.g. pH strips) or a pH meter is used.

10 REACTIONS BETWEEN IONS

10.1Interaction between ions

When we combine two strong electrolytes (AB and CD) we obtain a reaction mixture that contains the following ions:

A^+ B^- C^+ D^-

Two facts can occur:

ÿ or there is no interaction between the ions. They then continue to move freely next to each other. We propose this as follows:

AB + CD ÿ $A^+ + B^- + C^+ + D^-$

ÿ or there is interaction between certain ions. For example, suppose that the ions A^+ and D^- recombine to form a weak electrolyte AD. The representation of this is:

AB + CD ÿ AD + $B^- + C^+$

There are several possibilities for the weak electrolyte AD: ÿ either it is a precipitate (precipitation reactions) ÿ or it is a sparingly soluble gas (gas evolution reactions). ÿ or it is water (neutralization reactions)

10.2 Precipitation reactions

In a **precipitation reaction ,** ions recombine to form a sparingly soluble, solid compound. We call this a **precipitation.**
In a reaction equation, a precipitate is indicated by a downward pointing arrow (or by placing (v) after the substance).

example: A $BaCl_2$ solution contains Ba^{2+} ions and Cl^- ions.
A K_2SO_4 solution contains K^+ ions and SO_4 - ions.
After combining, the following reaction takes place:
$Ba^{2+} + 2\ K^+ + 2\ Cl^- + SO_4^{\ 2-}$ ÿ $BaSO_4$ÿ + $2\ K^+ + 2\ Cl^-$

In the **essential ion reaction,** only the ions that form a precipitate are noted
forms:

$Ba^{2+} + SO_4^{\ 2-}$ ÿ $BaSO_4$ÿ

The **substance reaction equation** is:
$BaCl_2$ (opl) + K_2SO_4 (opl) ÿ $BaSO_4$ÿ (v) + 2 KCl (opl)

You can look up whether a substance is soluble or not in a solubility table (see p. 23).

10.3Gas evolution reactions

In a gas evolution reaction, certain ions recombine to form a sparingly soluble, gaseous reaction product.

In a reaction equation, the formation of an escaping gas is indicated by an upward arrow.

example: combination of a Na2CO3 solution (dissociates) and an HCl solution (ionizes):

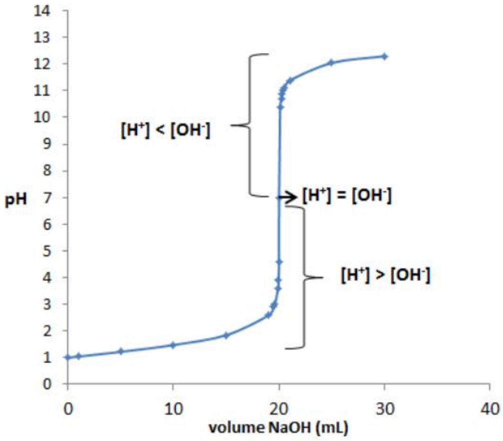
2 Na+ + CO3 2- + 2 H+ + 2 Cl- ÿ 2 Na+ + 2 Cl- + H2O + CO2ÿ

The essential ionic reaction is then: 2 H+ + CO3 2- ÿ H2O + CO2ÿ

The substance reaction equation is then:

2 HCl (opl)+ Na2CO3 (opl)ÿ H2O (vl) + CO2ÿ (g) + 2 NaCl (opl)

10.4Neutralization reactions



A **neutralization reaction** is a reaction in which H+ ions of an acid react with OH- ions of a base.
At the **equivalence point ,** all H+ ions have reacted with all OH- ions. The solution has become **neutral .** The solution here has a pH value of 7.
The equivalence point can be determined via a titration: this is the addition of two substances drop by drop (using a burette) in the presence of an indicator.
Titrations are performed to determine unknown concentrations of solutions.

Compounds	all **easily soluble**	**poorly soluble**
compounds with Na1+ compounds with K1+	at	**Neerslag¤** / **Kleur¤** AgCl¶ / Wit¶ AgBr¶ / Geel¶ AgI¶ / Geel¶ BaSO4¶ / Wit¶ Cu(OH)2¶ / Blauw¶ PbI2¶ / Ge¶ PbS¶ / Zwart¶ CuS¶ / Zwart¶ Fe(OH)3¤ / bruin¤
Salting of:		
Ammonium (NH4 1+) Nitraten (NO3 1-) Bromiden (Br1-) Chloriden (Cl1-) Iodide (I1-) Sulphates (SO4 2-) Sulfiden (S2-) Phosphates (PO4 3-) Carbonaten (CO3 2-)	at at all except ÿ all except ÿ all except ÿ all except ÿ Na1+, K1+ NH4 1+, Mg2+, Ba2+, Ca2+ Na1+, K1+, NH4 1+ Na1+, K1+, NH4 1+	 Ag1+ (Hg1+, Pb2+: hard) Ag1+ (Hg1+, Pb2+) Ag1+ (Hg1+, Hg2+ Pb2+) Ba2+ (Pb2+, Ca2+: matig) all andere everything else everything else
Hydroxiden (OH1-)	Group 1, more limited for group 2 other groups From the table	

you can deduce that all compounds with Na+ K+ or NO3 - are always well soluble.

11 REDOX REACTIONS

11.1The concepts of oxidation and reduction

In a redox reaction, electrons are always released by one element, allowing another element to absorb these electrons.

OXIDATION = *element releases e-*
REDUCTION = *element absorbs e-*

example. Mg + Cl2 => MgCl2 (white solid)

which element does e- absorb? *Cl2* which element gives off e- ? *Mg*

11.2The Oxidation Number (see also 5.3)

A redox reaction is always accompanied by a change in the oxidation number of these elements.

A correct definition of **oxidation number** is:
The oxidation number indicates the number of electrons that an atom absorbs/releases when it forms a bond with another atom (with ionic bonds there is a real electron transfer, while with covalent bonds we speak of a fictitious transfer).

OXIDATION NUMBER = charge that each atom gets or would get when it is a compound forms with another atom (group)

rules for determining oxidation numbers: H usually +I and O usually –II

H3PO4 *3*+I + x + 4*-II = 0 => x = +V*
SO4 2- *x + 4*-II = -2 => x = + VI*

OXIDATION = *increase in oxidation number of an atom (due to e - release; more + charge)*
REDUCTION = *reduction in oxidation number of an atom (due to e - absorption; more - charge)*

2 Mg + O2 ÿ 2 MgO
0 *0* *+II -II*

per element Mg worden is *2* e - *issued* (so *4* in total)
per element O worden is *2* e *included* (so *4* in total)

total number of e- issued = total number of e- withdrawn (e - balance in equilibrium)

REDOX REACTION = *reaction involving change in OG of elements.*
OXIDATOR = *element that* absorbs *e- (oxidizes but reduces itself)*
REDUCER = *element that* releases *e- (reduces but oxidizes itself)*

example: CuO+ H_2 => With + H_2O

which elements change from OG (and therefore also from e- number)?

+II 0 0 +I

which element oxidizes and which element reduces?

Cu reduces (from +II to 0 = decrease in OG, 1 times 2 e- included)

H oxidizes (from 0 to +I = increase in OG, 2 times 1 e- released)

what is the reducing agent and what is the oxidizing agent?

Cu is the oxidizer, H is the reducer

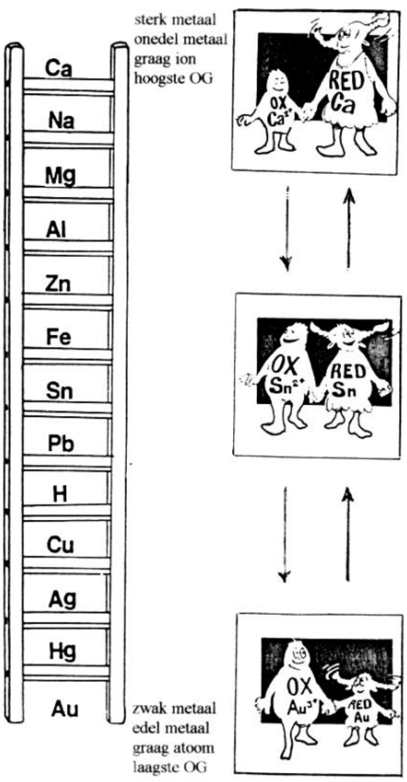
IN SUMMARY

Reducing agent or reducing substance	Oxidizer or oxidizing substance
is oxidized and releases e-	is reduced picks up e-mail
OG increases reduces the oxidizer	AND everything oxidizes the reducing agent

11.3Relative oxidizer and reducing agent strength

Metals can be arranged in an electrochemical **voltage series.**

The order in this series is determined by comparing how easily a metal atom gives up electrons to form positive ions in the solution. At the top of the series are the metals that easily donate electrons and thus easily form positive ions. These metals are called *strong or base* metals. The lower on the ladder, the less the metal atoms tend to donate electrons, the more difficult they are to form positive ions or the easier they are to survive as a metal atom. These metals are called *weak or noble* metals.



The figures next to the ladder with the voltage series of the metals are very meaningful.

The ion Ca^{2+} will have difficulty absorbing electrons from other metal atoms. **Ca^{2+} is the weakest *oxidizer*** depicted here. Conversely, a Ca atom will readily donate electrons to ions of all other metals depicted. **Ca is the strongest *reducing agent.***

The ion Au^{3+} will easily accept electrons from other metal atoms. **Au^{3+} is the strongest *oxidizer*** depicted here.

Conversely, an Au atom cannot donate electrons to ions of all other metals depicted. **Au is the weakest *reducer..***

12 CHEMICAL CALCULATIONS

12.1Amount of matter - mass

The atomic mass unit or 'unit' is taken as 1/12 of the mass of a 12C atom.

1 u = 1,66.10-24 g of 1,66 10-27 kg

The relative atomic mass Ar is an unnamed number that expresses how much the mass of an atom is greater than the mass of 1/12 part of a 12C atom.
The average relative atomic mass of each element is listed in the periodic table.

The relative molecular mass Mr is an unnamed number that expresses how much the mass of a molecule is greater than the mass of 1/12 part of a 12C atom. The relative molecular mass is the sum of the relative atomic masses of all atoms that make up that molecule.

A mole is a quantity containing NA particles. **NA = 6.02.1023 = Avogadro's number.**

One mole of atoms/molecules has a mass that is as many grams as the relative atomic mass/molecular mass indicates. This mass becomes the **molar mass *M***
named.

molar mass of an atom: ***M* = Ar g/mol**
example: M(Na) = 23.0 g/mol

molar mass of a molecule: ***M* = Mr g/mol**
example: M(Ca(OH)2 = 40.0 + 2*16.0 + 2*1.0 = 74.0 grams/mol

remark:
We round molar masses to the nearest 0.1!

We can represent an amount of matter by:
ÿ a number of particles *(N)*
ÿ a number of moles *(n)*
ÿ the mass corresponding to that number of particles or moles *(m)*

$$\text{PASTA (in g)} \underset{m = n\,.\,M}{\overset{n = m/M}{\rightleftarrows}} \text{NUMBER MOL} \underset{n = N/\,NA}{\overset{N = n\,.\,\text{AND}}{\rightleftarrows}} \text{NUMBER PARTICLES}$$

12.2Concentration of a solution

The concentration *c* is the ratio of the **amount of solute** to the **amount of solution.** This can be expressed in g/L or mol/L.

ÿ mass concentration cmass: number of grams of solute per liter of solution: = —

[cmassa] = g/L

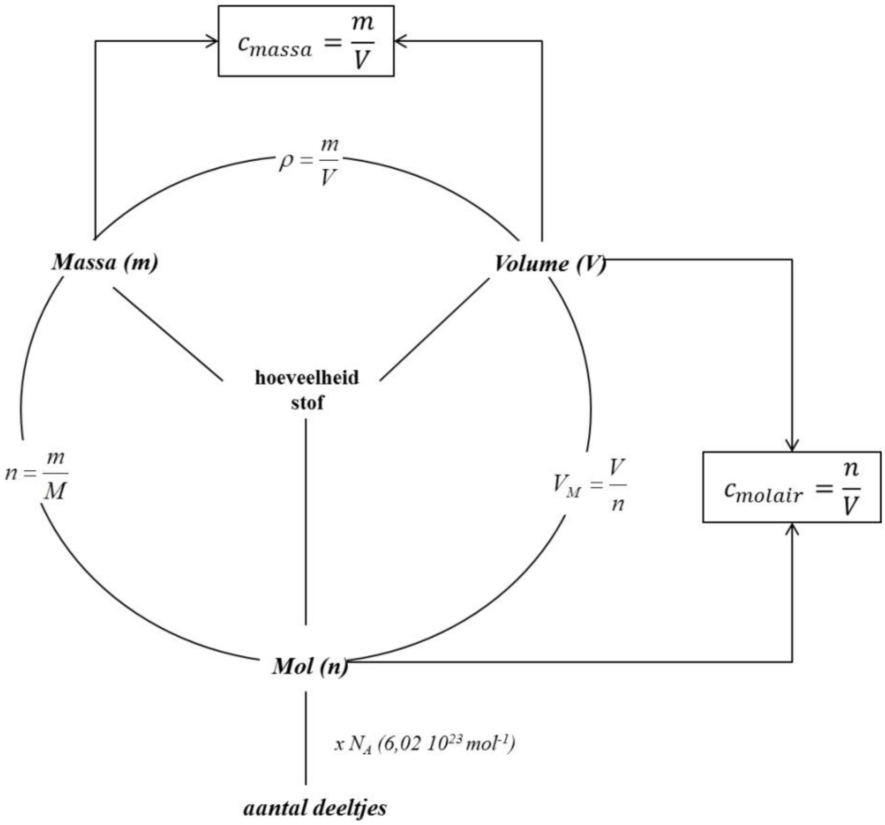
ÿ **molar concentration cmolar: number of moles of solute per liter of solution:**

= —

[cmolair] = mol/L

Notation: a solution of one liter containing 0.71 mol NaOH:
the molar concentration is 0.71 mol/L = 0.71 mol L-1

Conversion table:

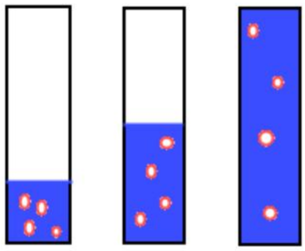


In the figure above, molar gas volume (VM) is *mentioned but not yet seen.*

When two solutions of the same substance but with a different concentration are added together, the new concentration can be calculated using the following formula:

1. 1 + 2. 2 = 1 + 2

When a solution is diluted, the number of moles *(n)* in the different solutions remains the same, but the concentration of the solution changes as shown in the figure.



Through the **general dilution law** we can find an unknown calculate concentration or an unknown volume:

cvoor . Vvoor = cna . Vna

n1 = n2 = n3

c1 > c2 > c3