BRIEF HISTORY of CHEMISTRY

Modern chemistry is about 250 years old as a science (dates back to the end of 18th century)

Ancient roots

common trades:

- **mining and metalworking, making alloys**
- **glass making**
- **food processing (wine and beer fermentation, vinegar production)**
- **soap making**

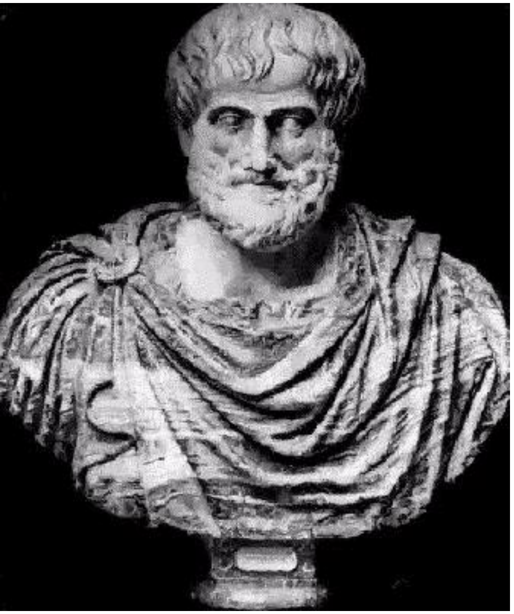
The word '*chemistry*' comes from alchemy, but its origin is ambiguous: it may have arabic, greek or egyptian roots.

Greek roots

Democritus (460–370 BC): ***philosopher***

atomic theory (atomism): ‚atomos' = indivisible

- **everything is composed of "atoms"**
- **atoms are physically indivisible**
- **there are different kinds of atoms, which differ in shape and size.**



Aristotle (384–322 BC): ***philosopher and scientist***

fundamental elements and properties:

Element	Hot/Cold	Wet/Dry	Modern state of matter
Earth	Cold	Dry	Solid
Water	Cold	Wet	Liquid
Air	Hot	Wet	Gas
Fire	Hot	Dry	Plasma
Aether	divine substance of the heavenly spheres, stars and planets		

Middle Ages

chemical medicine (or iatrochemistry)

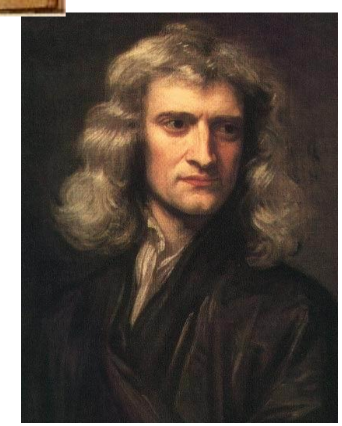
- seeks to provide chemical solutions to diseases and medical ailments (finding or making medicines)
- physiological effects of certain substances

alchemy:

- set of practices including metallurgy, philosophy, astrology, astronomy, mysticism, medicine
- quest for the philosopher's stone with which one can transform ordinary metals into gold
- by performing experiments and recording the results, alchemists set the stage for modern chemistry
- Roger Bacon (1214 – 1294), Nicolas Flamel (1340 – 1418), Tycho Brahe (1546–1601), Robert Boyle (1627–1691) etc.

Paracelsus (1493-1541)

Jabir ibn Hayyan, Geber (721-815)



Sir Isaac Newton (1643-1727)

Chemical element, 1661

Robert Boyle
(1627-1691)

An element is a fundamental substance that cannot be chemically changed or broken down into anything simpler.

MODERN DEFINITION: a species of atoms having the same number of protons in their atomic nuclei

Georg Ernst Stahl (1659-1734)

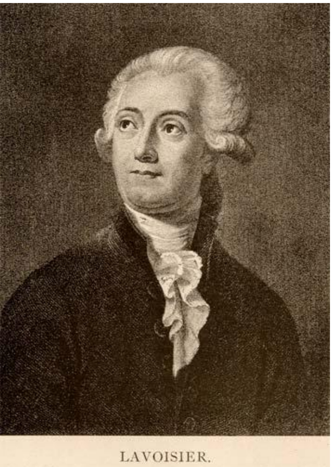
Phlogiston theory, 1703

- theory attempted to explain processes such as combustion and making metals
- phlogiston: fire-like element, contained within combustible bodies and released during combustion

  e.g.: wood = ash + phlogiston

- problems: some metals gain mass when they burn

Chemistry as a quantitative science (modern chemistry)



Antoine Laurent Lavoisier (1743-1794)

- discovered the role of oxygen in combustion
- recognized and named oxygen (1778) and hydrogen (1783)
- helped to construct the metric system,
- helped to reform chemical nomenclature.
- discovered that, although matter may change its form or shape, its mass always remains the same.

Mikhail Vasilyevich Lomonosov (1711-1765)

- together with Lavoisier, is regarded as the one who discovered the law of mass conservation.
- contributed to the formulation of the kinetic theory of gases
- published a catalogue of over 3,000 minerals
- demonstrated the organic origin of soil, peat, coal, petroleum and amber

Progress from alchemy to modern chemistry

LAW OF MASS CONSERVATION

Mass is neither created nor destroyed in chemical reactions.

TODAY: Not exactly true

Einstein's mass-energy equivalence equation

$$\Delta E = \Delta m c^2$$

~100 kJ/mol ~10^{-12} kg/mol = 1 ng/mol

LAW OF DEFINITE PROPORTIONS (1799)

Different samples of a pure chemical substance always contain the same proportion of elements by mass.

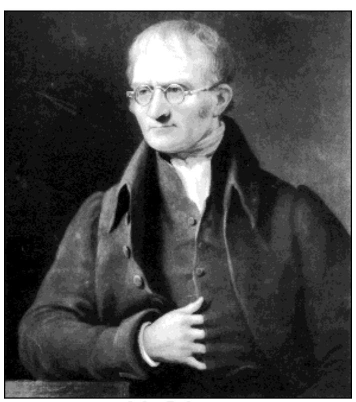
The composition of a given compound does not depend on its source and method of preparation

$$1.000\ \text{g Cu} \xrightarrow{400\ °\text{C}} 1.252\ \text{g CuO}$$

$$1.000\ \text{g Cu} \xrightarrow{HNO_3} Cu(NO_3)_2 \xrightarrow{400\ °\text{C}} 1.252\ \text{g CuO}$$

Joseph Louis Proust (1754-1826)

LAW OF MULTIPLE PROPORTIONS (1803)



John Dalton

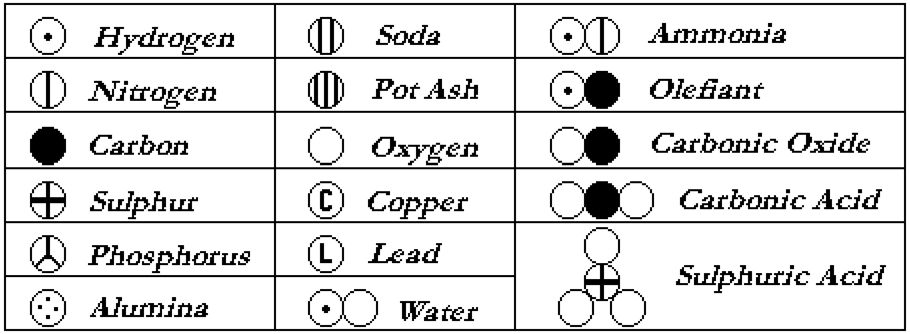
John Dalton (1766-1844)

Elements can combine in different ways to form different substances, whose mass ratios are small whole-number multiples of each other.

compound	m/m% Cu	m/m% O	m_{Cu}/m_O	ratio
Cu_2O	88.8	11.2	7.94	2
CuO	79.9	20.1	3.97	1

Atomic theory of Dalton (1808)

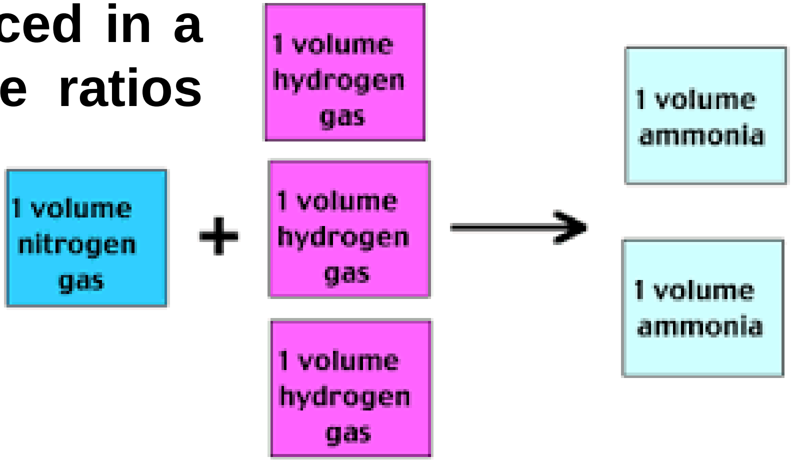
- Elements are made of tiny particles called *atoms*.
- Each element is characterized by the mass of its atoms. Atoms of the same element have the same mass, but atoms of different elements have different masses.
- Chemical combination of elements to make different substances occurs when atoms join together in small whole-number ratios.
- Chemical reactions only rearrange the way that atoms are combined; the atoms themselves are unchanged.



LAW OF COMBINING GAS VOLUMES (1808)

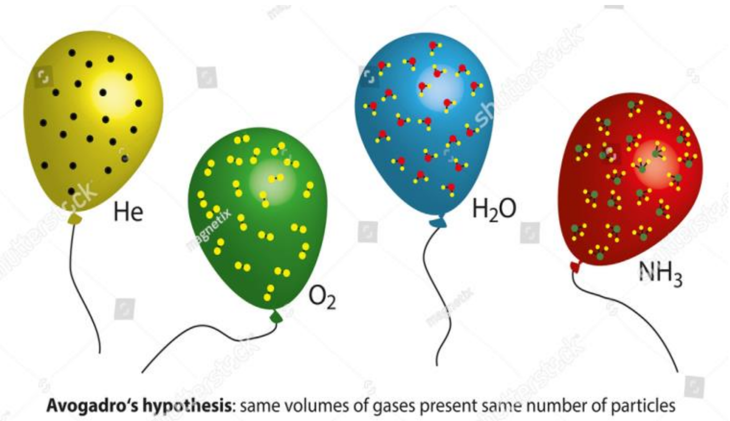
Joseph Louis Gay-Lussac (1778-1850)

Gases at constant temperature and pressure react or are produced in a chemical reaction in volume ratios of the small whole numbers.



$$N_2 + 3H_2 = 2NH_3$$

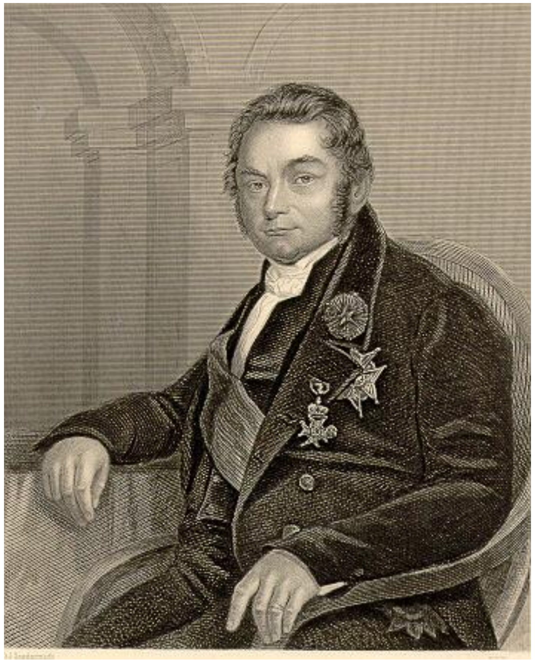
AVOGADRO'S LAW (1811)



Identical volumes of different gases contain the same number of molecules at the same temperature and pressure.

Amedeo Avogadro (1776-1856)

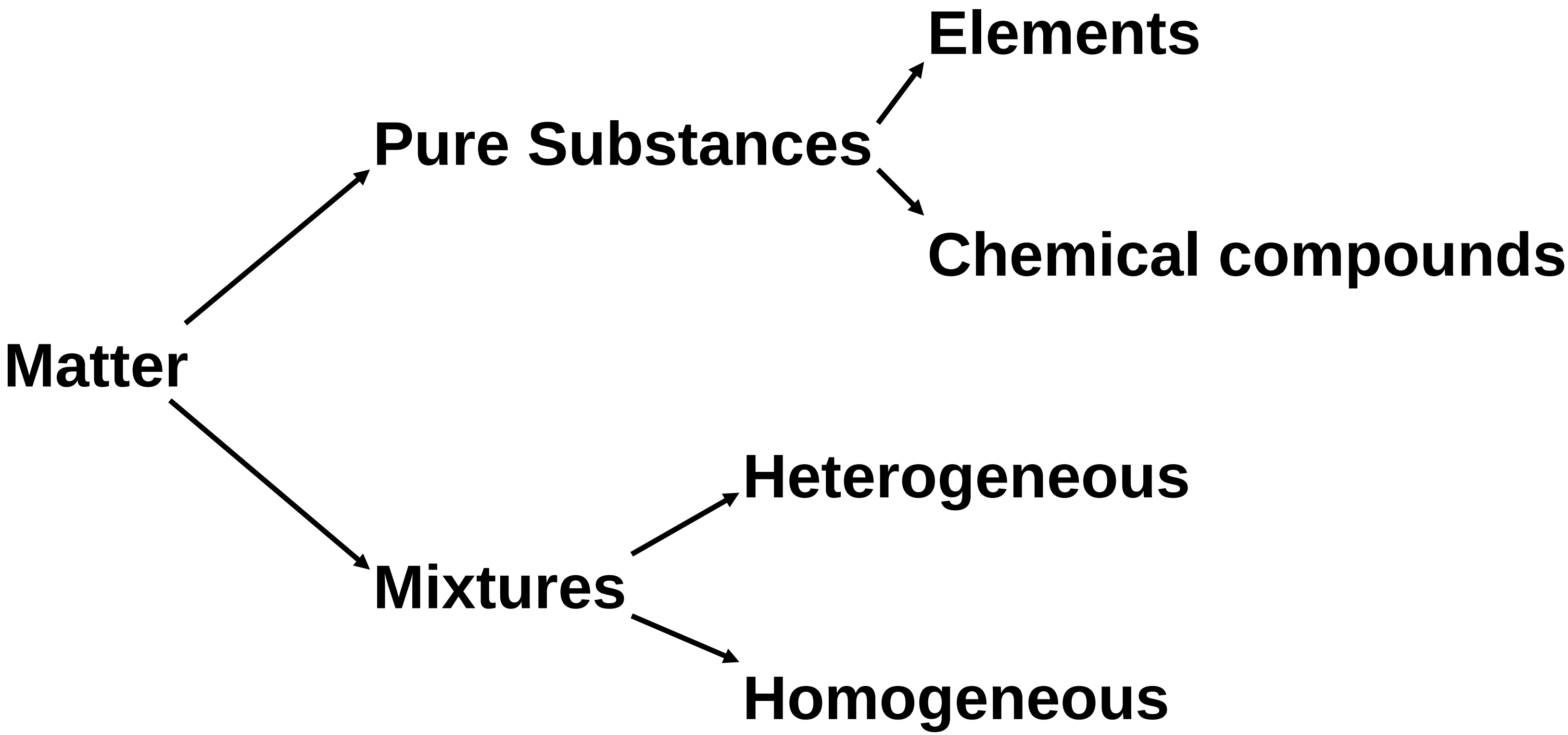
Chemical symbols of elements



Jöns Jakob Berzelius (1779-1848)

- Study of 47 elements
- Determination of atomic masses
- 1813: suggestion of using letters
- Most of them are still used
- Some exceptions

Name	Berzilius	Now
Chromium	Ch	Cr
Manganese	Ma	Mn
Magnesium	Ms	Mg
Iridium	I	Ir
Tungsten	Tn	W
Potassium	Po	K
Sodium	So	Na


Matter
Pure Substances
Elements
Chemical compounds
Mixtures
Heterogeneous
Homogeneous

Mixtures

simple blend of two or more substances added together in some random proportion without chemically changing the individual substances themselves.

Heterogeneous: the mixing is not uniform, the mixture has regions of different composition

e. g. two unmiscible liquids: water + oil

soil, pizza, cereal in milk

Homogeneous: the mixing is uniform

e.g. gaseous: dust-free air

liquid: water-ethanol mixture

solid: certain alloys

Compounds

pure substance that is formed when atoms of two or more different elements combine and create a new material with properties completely unlike those of its constituent elements.

Notation of pure substances

element → chemical symbol

compound → chemical formula

$NaCl$ for sodium chloride or common salt

formula	chemical name systematic name	common name trivial name

Formulas: different types for different purposes

Formulas

empirical formula: shows the ratios of atoms in a compound

NaCl CH_2O

molecular formula: shows the actual number of atoms in a molecule

----- CH_2O $(CH_2O)_6 = C_6H_{12}O_6$

structural formula: shows how the atoms are connected within a molecule

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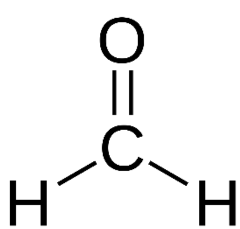
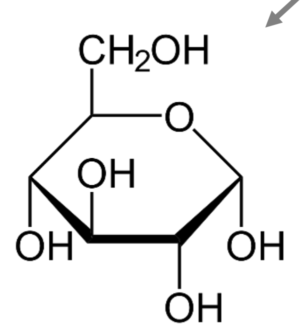
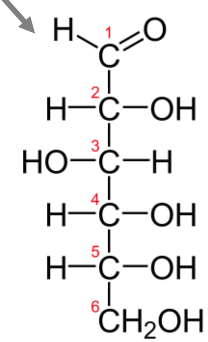






NAMING COMPOUNDS

only inorganic compounds, organic compounds will be covered as part of the Organic Chemistry course

NOMENCLATURE = system of names, naming rules

present chemical nomenclature:

discussed and approved by *IUPAC*

International Union of Pure and Applied Chemistry

BINARY IONIC COMPOUNDS

name = positive ion + negative ion

positive ion = name of the element

negative ion = name of the element + suffix -ide

KCl: potassium chloride

K always forms K^+ ions

LiF: lithium fluoride

Li always forms Li^+ ions

$BaCl_2$: barium chloride

Ba always forms Ba^{2+} ions

$AlBr_3$: aluminum bromide

Al always forms Al^{3+} ions

$CrCl_3$: chromium(III) chloride

Cr can form both Cr^{2+} and Cr^{3+} ions: $CrCl_2$: chromium(II) chloride

PbS: lead(II) sulfide

PbS_2: lead(IV) sulfide

$FeCl_2$: iron(II) chloride

$FeCl_3$: iron(III) chloride

$SnCl_2$: tin(II) chloride

$SnCl_4$: tin(IV) chloride

Cu_2O: copper(I) oxide

CuO: copper(II) oxide

BINARY MOLECULAR COMPOUNDS

***name = more positive part + more negative part* + suffix –ide**

use of Greek numbers in names:

1 --- (mono)	**3 tri**	**5 penta**	**7 hepta**
2 di	**4 tetra**	**6 hexa**	**8 octa**

HF: hydrogen fluoride

CO: carbon monoxide

CO_2: carbon dioxide

$AsCl_3$: arsenic trichloride

$SeBr_4$: selenium tetrabromide

PCl_5: phosphorus pentachloride

SF_6: sulfur hexafluoride

IF_7: iodine heptafluoride

N_2O: dinitrogen monoxide

NO: nitrogen monoxide

N_2O_3: dinitrogen trioxide

NO_2: nitrogen dioxide

N_2O_4: dinitrogen tetroxide

N_2O_5: dinitrogen pentoxide

Exception

P_2O_5: phosphorus pentoxide (instead of diphosphorus pentoxide)

Some metal compounds (often oxides) are molecular and both method can be used

OsO_4: osmium tetroxide *or* osmium(VIII) oxide

CrO_3: chromium trioxide or chromium(VI) oxide

CrO_2: chromium dioxide or chromium (IV) oxide

PbO_2: lead dioxide or lead(IV) oxide

MnO_2: manganese dioxide or manganese(IV) oxide

Note:

V_2O_5: <u>vanadium pentoxide</u> (instead of divanadium pentoxide) or vanadium(V) oxide

POLYATOMIC IONS

names must be learnt e.g.

CO_3^{2-} carbonate ion

SO_4^{2-} sulfate ion

NO_3^{-} nitrate ion

PO_4^{3-} phosphate ion

NH_4^{+} ammonium ion

$(NH_4)_2SO_4$ ammonium sulfate

$Fe(NO_3)_2$ iron(II) nitrate

$Fe(NO_3)_3$ iron(III) nitrate

$Hg(NO_3)_2$ mercury(II) nitrate

KCN potassium cyanide (CN^- cyanide ion)

Chemical equations

does show - identity of reactants and products

- stoichiometry (quantitative relationship between reacting or produced substances e.g. mass, amount of substance)

- direction of spontaneous reaction

$$2H_2 + O_2 \rightarrow 2H_2O \text{ or } 2H_2 + O_2 = 2H_2O$$

may show - phase of reactants and products

HCl (g) $\Rightarrow$ in gas phase CH_4O (l) $\Rightarrow$ in liquid phase

SiO_2 (s) $\Rightarrow$ in solid phase

$NaCl$ (aq) $\Rightarrow$ in aqueous phase = dissolved in water

$$CaCO_3\,(s) + 2CH_3COOH\,(l) \rightarrow Ca(CH_3COO)_2\,(aq) + H_2O\,(l) + CO_2\,(g)$$

- energy change of the reaction

$$2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(g)\ \Delta E = -484\ \text{kJ/mol}$$

does not usually show

- time needed for completion
- conditions (e.g. temperature)

molecular equation:

$$AgNO_3 + NaCl \rightarrow \underline{AgCl} + NaNO_3$$

ionic equation: electrolytes are written as dissociated ions

$$Ag^+ + \cancel{NO_3^-} + \cancel{Na^+} + Cl^- \rightarrow \underline{AgCl} + \cancel{Na^+} + \cancel{NO_3^-}$$

net ionic equation:

$$Ag^+ + Cl^- \rightarrow \underline{AgCl}$$

spectator ions: Na^+, NO_3^-

Balancing chemical equations

conservation of mass and conservation of charge

$$3CaCl_2 + 2Na_3PO_4 \rightarrow Ca_3(PO_4)_2 + 6NaCl$$

Ca:	3	= 3
Cl:	3×2 = 6	= 6
Na:	2×3 = 6	= 6
P:	2	= 2
O:	2×4 = 8	= 2×4 = 8
charge:	0	= 0

Balancing chemical equations

$$2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$$

S:	$2 \times 2 = 4$	=	4
O:	$2 \times 3 = 6$	=	6
I:	2	=	$2 \times 1 = 2$
charge:	$2 \times -2 = -4$	=	$-2+2 \times -1 = -4$

Balancing chemical equations

Step 1: Write reactants and products

$$C_3H_8 + O_2 \rightarrow CO_2 + H_2O$$

Step 2: Find one atom that occurs only in one substance on both sides

$$C_3H_8 + O_2 \rightarrow CO_2 + H_2O$$

Step 3: Find coefficients to balance this atom

$$1C_3H_8 + O_2 \rightarrow 3CO_2 + H_2O$$

Step 4: Find another unbalanced atom which occurs in only one substance

$$1C_3H_8 + O_2 \rightarrow 3CO_2 + H_2O$$

Step 5: Find coefficient to balance this atom

$$1C_3H_8 + O_2 \rightarrow 3CO_2 + 4H_2O$$

Step 6: Repeat steps 4-5 until you have balanced all atoms

$$1C_3H_8 + O_2 \rightarrow 3CO_2 + 4H_2O$$

$$1C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

Step 7: Make sure the coefficients are reduced to the smallest whole number values

$$1C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

Step 8: Check the balanced equation

$$1C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

3C, 8H, 10O 3C, 10O, 8H

GROUPS OF CHEMICAL REACTIONS

Based on the number of reagents and products

(mainly in organic chemistry)

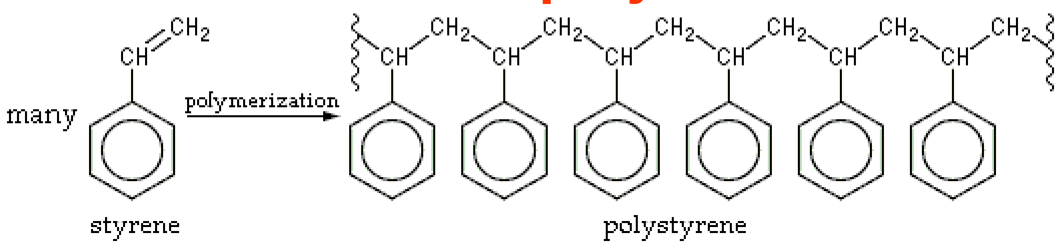
- **isomerization:** $NH_4OCN \rightarrow H_2N–CO–NH_2$
- **addition:** $CH_2{=}CH_2 + Cl_2 \rightarrow CH_2Cl–CH_2Cl$

  special cases:

  dimerization: $2C_5H_6 \rightarrow C_{10}H_{12}$

  trimerization: $3C_2H_2 \rightarrow C_6H_6$

  polymerization:



- **elimination/decomposition:** $NH_4Cl \rightarrow NH_3 + HCl$
- **substitution:** $C_2H_5Cl + NaOCH_3 \rightarrow C_2H_5–O–CH_3 + NaCl$

GROUPS OF CHEMICAL REACTIONS

Based on the observable change

- gas formation: $Zn + 2HCl \rightarrow ZnCl_2 + H_2$

$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$

- color change: $2KBr + Cl_2 \rightarrow 2KCl + Br_2$

KBr: colorless aqueous solution

Br_2: brownish is water

- precipitation: $ZnCl_2 + H_2S \rightarrow \underline{ZnS} + 2HCl$

$AgNO_3 + NaCl \rightarrow \underline{AgCl} + NaNO_3$

- combustion: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

GROUPS OF CHEMICAL REACTIONS

Based on the transferred species

electron transfer occurs: REDOX REACTION

$$2Mg\ (s) + O_2\ (g) \rightarrow 2MgO\ (s)$$

no electron transfer occurs: ACID-BASE REACTION

$$NaOH\ (aq) + HCl\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l)$$

REDOX REACTIONS

oxidation: the loss of one ore more electrons

reduction: the gain of one ore more electrons

reduction

$$Fe^{2+} + 2e^{-} \rightarrow Fe$$

$$Fe^{3+} + e^{-} \rightarrow Fe^{2+}$$

$$Fe^{3+} + 3e^{-} \rightarrow Fe$$

$$Cl_2 + 2e^{-} \rightarrow 2Cl^{-}$$

$$O_2 + 4e^{-} \rightarrow 2O^{2-}$$

oxidation

Earlier definitions: still used widely in organic chemistry

oxidation: addition of oxygen (O) or removal of hydrogen (H)

reduction: removal of oxygen (O) or addition of hydrogen (H)

$$CH_4 \xrightarrow{+O} CH_4O\ (CH_3OH) \xrightarrow{-2H} CH_2O \xrightarrow{+O} CH_2O_2\ (HCOOH) \xrightarrow{-2H} CO_2$$

Oxidation number (oxidation state)

a formal number assigned to every atom in a compound or element, which indicates whether the atom is neutral, electron-rich or electron-poor

– useful tool for classifying compounds and balancing equations

– does not necessarily imply actual charges on atoms

ELECTRONEGATIVITY (EN):

a dimensionless quantity showing the ability of an atom in a molecule to attract the shared electrons in a covalent bond

F: 4.0	**O: 3.5**	**N: 3.0**	**Cl: 3.0**
H: 2.1	**Al: 1.5**	**K: 0.8**	**Fr: 0.7**

Rules for assigning oxidation numbers

General rules

1. An atom in its elemental state has an oxidation number of 0.

Na, H_2, Br_2, F_2, O_2, Ne, S_8, P_4 oxidation number 0

2. An atom in a monatomic ion has an oxidation number identical to its charge.

Na^+: +1 Ca^{2+}: +2 Al^{3+}: +3 Cl^-: -1 O^{2-}: -2

3. The sum of the oxidation numbers is 0 for a neutral compound

H_2SO_3 H: +1 S: +4 O: -2 $KMnO_4$ K: +1 Mn:+7 O: -2

4. The sum of the oxidation numbers equals to the net charge for a polyatomic ion.

NH_4^+ N: -3 H: +1 CO_3^{2-} C: +4 O: -2

5. In a heteronuclear covalent bond, the bonding electron pair is assigned to the more electronegative atom.

H–Cl H: +1 Cl: -1 H–O–H H: +1 O: -2

Specific rules

1. Fluorine always has an oxidation number of -1 in its compounds.

XeF_2 Xe: +2 F: -1 SF_6 S: +6 F: -1

2. Alkali metals (main group I) always have an oxidation number of +1, alkaline earth metals (main group II) +2 in their compounds.

LiH Li: +1 H: -1 BaO_2 Ba: +2 O: -1

3. Aluminum usually has an oxidation number of +3.

$LiAlH_4$ Li: +1 Al: +3 H: -1 $AlBr_3$ Al: +3 Br: -1

Zinc, cadmium usually have an oxidation number of +2; silver usually has an oxidation number of +1.

$ZnCl_2$ Zn: +2 Cl: -1 CdI_2 Cd: +2 I: -1 Ag_2S Ag: +1 S: -2

4. Hydrogen usually has an oxidation number of +1 except for the metal hydrides (when it forms a compound with an element of lower electronegativity)

H_2S H: +1 S: -2   BUT CaH_2 Ca: +2 H: -1

5. Oxygen usually has an oxidation number of -2.

XeO_2 Xe: +4 O: -2 **SO_2 S: +4 O: -2**

Notable exceptions: **oxygen fluorides OF_2 O: +2 F: -1**

peroxides: H_2O_2 H: +1 O: -1 **superoxides: KO_2 K: +1 O: -½**

6. Halogens have an oxidation number of -1 in all compounds not containing oxygen or another halogen.

$FeCl_3$ Fe: +3 Cl: -1 $FeBr_2$ Fe: +2 Br: -1

<u>Other</u> cases:

$NaClO_2$ Na: +1 Cl: +3 O: -2 **ClO_4^- Cl: +7 O: -2**

BrO_2 Br: +4 O: -2 **HClO H: +1 Cl: +1 O: -2**

BrCl Br: +1 Cl: -1 **ClF_3 Cl: +3 F: -1** **IF_5 I: +5 F: -1**

There are cases when more then one assignment can be used without problems. **Fe_3O_4** **Fe: +8/3 O: -2**

Fe_2: +3 Fe_1: +2 O: -2

REDUCING AGENT:

- causes reduction
- loses one or more electrons
- undergoes oxidation
- oxidation number of atom increases

OXIDIZING AGENT:

- causes oxidation
- gains one or more electrons
- undergoes reduction
- oxidation number of atom decreases

$$\overset{+1\ +7\ -2}{2KMnO_4} + \overset{+1\ \ +4\ -2}{5Na_2SO_3} + \overset{+1\ +6\ -2}{3H_2SO_4} \rightarrow$$

$$\overset{+2\ +6\ -2}{2MnSO_4} + \overset{+1\ \ +6\ -2}{5Na_2SO_4} + \overset{+1\ +6\ -2}{1K_2SO_4} + \overset{+1\ -2}{3H_2O}$$

Mn: +7 → +2 decrease, oxidizing agent

S in Na_2SO_3: +4 → +6 increase, reducing agent

Mn: -5 change **2 mol needed**

S in Na_2SO_3: +2 change **5 mol needed**

$$2\overset{+1}{K}\overset{+7}{Mn}\overset{-2}{O_4} + 3\overset{+1}{Na_2}\overset{+4}{S}\overset{-2}{O_3} + 1\overset{+1}{H_2}\overset{-2}{O} \rightarrow$$

$$2\overset{+4}{Mn}\overset{-2}{O_2} + 3\overset{+1}{Na_2}\overset{+6}{S}\overset{-2}{O_4} + 2\overset{+1}{K}\overset{-2}{O}\overset{+1}{H}$$

Mn: +7 → +4 decrease, oxidizing agent

S: +4 → +6 increase, reducing agent

Mn: -3 change	**2 mol needed**
S: +2 change	**3 mol needed**

$$\overset{+1\ +7\ -2}{2KMnO_4} + \overset{+1\ +4\ -2}{1Na_2SO_3} + \overset{+1\ -2\ +1}{2KOH} \rightarrow$$

$$\overset{+1\ +6\ -2}{2K_2MnO_4} + \overset{+1\ +6\ -2}{1Na_2SO_4} + \overset{+1\ -2}{1H_2O}$$

Mn: +7 → +6 decrease, oxidizing agent

S: +4 → +6 increase, reducing agent

Mn: -1 change	**2 mol needed**
S : +2 change	**1 mol needed**

SIMPLE EXAMPLES OF REDOX REACTIONS I.

Combustion: burning of fuel by oxidation with oxygen in air. Fuel: gasoline, natural gas, fuel oil, wood, carbon

$$CH_4\ (g) + 2O_2\ (g) \rightarrow CO_2\ (g) + 2H_2O\ (l)$$

Bleaching: redox reactions to decolorize or lighten colored materials (hair, wood pulp, clothes)

Bleaching agent = tipically oxidizing agent

usually H_2O_2, $NaOCl$, Cl_2

Batteries: providing electrical current using spontaneous chemical reactions. Redox reactions are needed based on common and inexpensive materials

$$Zn\ (s) + 2MnO_2\ (s) + 2NH_4Cl\ (s) \rightarrow$$

$$ZnCl_2\ (aq) + Mn_2O_3\ (s) + 2NH_3\ (aq) + H_2O\ (l)$$

SIMPLE EXAMPLES OF REDOX REACTIONS II.

Metallurgy: the science of extracting and purifying metals from their ores, always based on redox reactions

$$ZnO\ (s) + C(s) \rightarrow Zn\ (s) + CO\ (g)$$

Corrosion: deterioration of a metal by oxidation, rusting, usually oxygen from air and moisture is needed

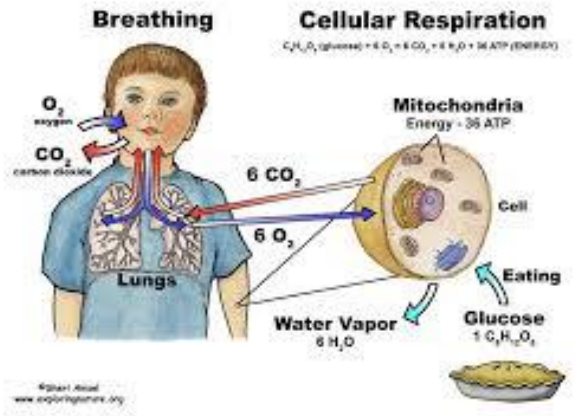
$$4Fe\ (s) + 3O_2\ (g) + 2H_2O\ (l) \rightarrow 2Fe_2O_3 \cdot H_2O\ (s)$$

iron rust

Respiration: the process of breathing and using oxygen for the biological redox reactions that provide the energy needed by living organisms.

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O + \text{biological energy}$$

glucose (a sugar)



Simple examples of acid-base reactions I.

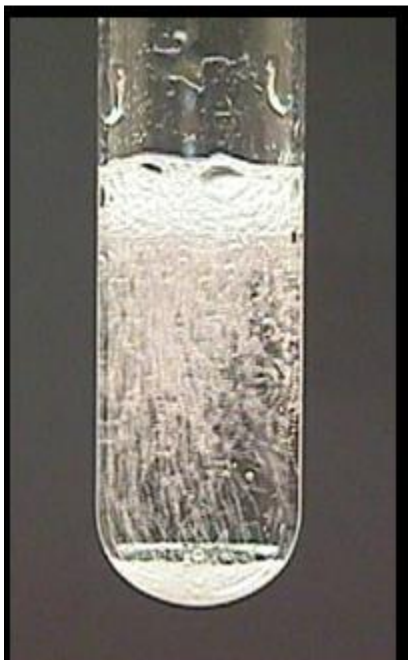
precipitation

$AgNO_3 + NaCl \rightarrow \underline{AgCl} + NaNO_3$

$Pb(NO_3)_2 + 2KI \rightarrow \underline{PbI_2} + 2KNO_3$

$FeCl_3 + 3NaOH \rightarrow \underline{Fe(OH)_3} + 3NaCl$

precipitate



gas formation

$FeS + 2HCl \rightarrow FeCl_2 + H_2S\ (g)$

$Na_2CO_3 + H_2SO_4 \rightarrow Na_2SO_4 + H_2O + CO_2\ (g)$

$NH_4Cl + NaOH \rightarrow NaCl + H_2O + NH_3\ (g)$

Simple examples of acid-base reactions II.

neutralization

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

$$CH_3COOH + NH_3 \rightarrow NH_4CH_3COO$$

$$2CsOH + H_2SO_4 \rightarrow Cs_2SO_4 + 2H_2O$$

complex formation

$$\underline{AgCl} + 2NH_3 \rightarrow [Ag(NH_3)_2]Cl$$

$$\underline{HgO} + 4KI + H_2O \rightarrow K_2[HgI_4] + 2KOH$$

$$Zn^{2+} + 4OH^{-} \rightarrow [Zn(OH)_4]^{2-}$$

Chemistry as a quantitative science

Starting from the second half of 18th century with the works of Lavoisier and Lomonosov

- **accurate measurements (mass, volume, temperature, pressure etc)**
- **discovery of gases, gas laws**
- **19th century: foundation of electrochemistry and thermochemistry**

Quantitative measurements

$$V = 330\ \mathrm{ml}$$

physical property	**number**	**unit**
(volume)	**(330)**	**(milliliter)**

Physical properties

extensive	⇔	**intensive**
related to size, e.g. amount of matter, volume, mass, energy		unrelated to size e. g. temperature, density, pressure, concentration
fundamental	⇔	**derived**
not a combination of other properties e.g. length, mass, time		a combination of other properties e.g. area, volume, density, velocity, acceleration

1960: Systéme Internationale d'Unites (SI)

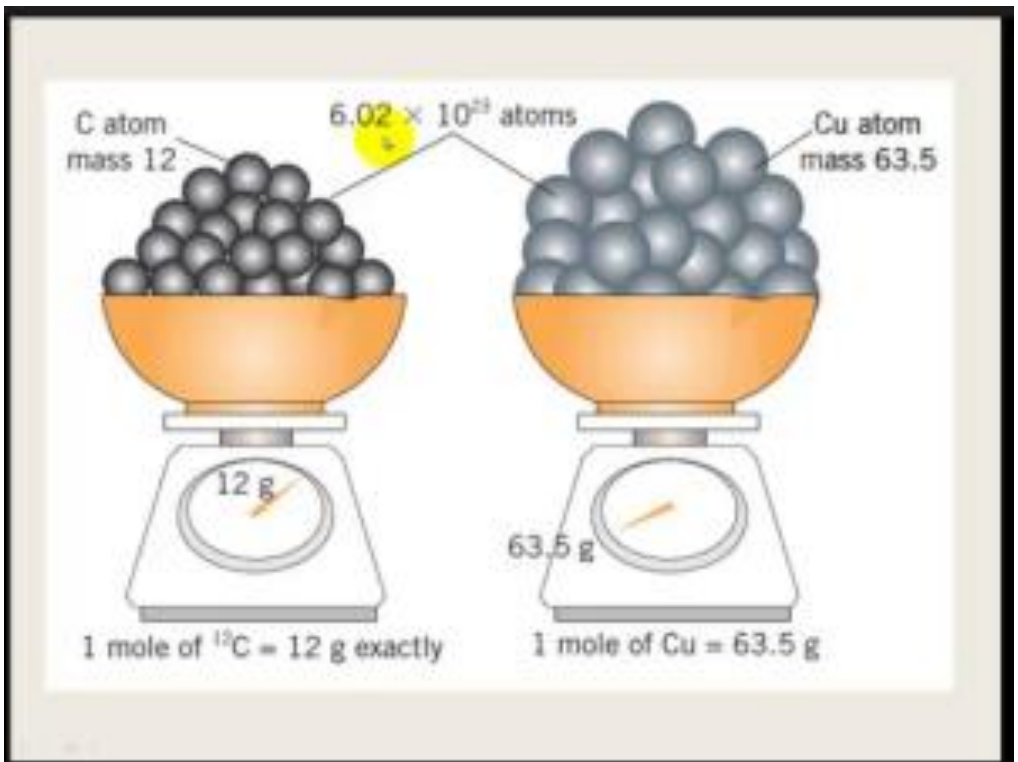
earlier systems: e.g. CGS (cm, g, s), later MKS (m, kg, s)

Fundamental quantities and units in SI

quantity	symbol	name of basic unit	symbol of unit	basic unit defined by	from 20/May/2019 the definition is based on
mass	m	kilogramm	kg	international standard kept in Paris	h (Planck constant)
length	l	meter	m	speed of light (c)	*unchanged*
temperature	T	kelvin	K	triple point of water	k_B (Boltzmann constant)
amount of substance	n	mole	mol	number of atoms in 12 grams of ^{12}C	N_A (Avogadro constant)
time	t	second	s	period of radiation between energy levels of ^{133}Cs	*unchanged*
electric current	I	ampere	A	current between two conductors under the defined conditions	e (elementary charge)
luminous intensity	I_V	candela	cd	luminous intensity of a defined light source	*unchanged*

Amount of substance: mole (mol)

Current definition: The amount of a chemical substance that contains as many representative particles, e.g., atoms, molecules, ions, electrons, or photons, as there are atoms in 12.0 grams of carbon-12.



The number of particles present in one mole of substance can be expressed by the use of the Avogadro constant

$$N_A = 6.022 \times 10^{23}\ \text{mol}^{-1}$$

New definition (from 20th May 2019): the amount of substance of exactly $6.02214076 \times 10^{23}$ elementary entities.

Prefixes in SI

precedes a basic unit to indicate a multiple or fraction of the unit

e.g. **1 nm = 10^{-9} meter**

1 kmol = 10^{3} mol

prefix	symbol	factor	prefix	symbol	factor
giga	G	10^{9}	deci	d	10^{-1}
mega	M	10^{6}	centi	c	10^{-2}
kilo	k	10^{3}	milli	m	10^{-3}
hecto	h	10^{2}	micro	μ	10^{-6}
deka	da	10^{1}	nano	n	10^{-9}

There are further prefixes denoting larger quantities (exa, peta, tera etc.) and smaller quantities (pico, femto, atto etc.).

Significant figures (s.f.)/significant digits

- digits that carry meaning contributing to the result of a measurement
- 25 ≠ 25.0 ≠ 25.00 ≠ 25.000 etc.
- all digits but the last are certain

  → the last is a best guess, usually having an error (± 1, 2..)

Identifying significant figures (s.f.)

- All digits other than 0 are significant
- Zeros in the middle of a number are always significant (540826 → 6 s.f.)
- Zeros at the beginning of a number are never significant (0.0058426 → 5 s.f.)
- Zeros at the end of the number and after the decimal point are always significant (1.560 → 4 s.f.)
- Zeros at the end of a number and before the decimal point may or may not be significant (15200 → 3 or 5 s.f.)

Significant figures in calculations

1. Multiplication or division: the answer cannot have more significant numbers than either of the original numbers

$11.78945 \text{ g} / 11.9 \text{ cm}^3 = 0.991 \text{ g/cm}^3$

2. Addition or subtraction: the answer cannot have more digits to the right of the decimal point than either of the original numbers

3.18 g + 0.01315 g = 3.19 g

Rounding numbers

5.664525

1. If the first digit you remove is less than 5, round by dropping it and all following digits (e.g. round down). (3 s.f.) $\rightarrow$ 5.66

2. If the first digit you remove is 6 or greater, round by adding 1 to the digit on the left (e.g. round up). (2 s.f.) $\rightarrow$ 5.7

3. If the first digit you remove is 5 and there are more nonzero digits following, round up. (4 s.f.) $\rightarrow$ 5.665

4. If the digit you remove is 5 with nothing following, round down. (6 s.f.) $\rightarrow$ 5.66452