

Glossary - Chemistry (Intro)

Chemistry: The science that studies the properties of substances and how substances react with one another.

Chemistry in contrast to physics is involved with the change of the outer electron-layers only, whereas physics involves the state of the nucleus as well.

C. **Energy:** Energy stored within the structural units of chemical substances [J], see thermochemistry.

C. **Equation:** An equation that uses chemical symbols to show what happens during a chemical reaction.

C. **Equilibrium:** A state in which the rates of the forward and reverse reactions are equal.

C. **Formula:** An expression showing the chemical composition of a compound in terms of the symbols for the atoms of the elements involved; see stoichiometry.

C. **Kinetics:** The area of chemistry concerned with the speeds, or rates, at which chemical reactions occur; see stoichiometry.

C. **Property:** Any property of a substance that cannot be studied without converting the substance into some other substances.

Electro C.: The branch of chemistry that deals with the use of chemical reactions to produce electricity, the relative strengths of oxidizing and reducing agents, and the use of electricity to produce chemical change.

Inorganic C.: The branch of chemistry that deals with compounds other than organic compounds.

Organic C.: The branch of chemistry that deals with carbon and usually hydrogen compounds, excluding carbohydrates.

Nuclear C.: The study of the structure of nuclei, of the changes this structure undergoes, and of the consequences of those changes for chemistry.

Thermo C.: The study of heat changes in chemical reactions.

Closed System: A system that allows the exchange of energy (usually in the form of heat) but not mass with its surrounding environment.

Cohesion: The intermolecular attraction between like molecules (see physics - matter).

Compound: 1) A specific combination of elements that can be separated into elements by using chemical techniques. 2) A substance consisting of atoms of two or more elements in a defined ratio.

Distillation: The separation of a mixture by making use of the different volatilities of its compounds.

Fractional D.: Separation of the components of a liquid mixture by repeated distillation, by making use of their differing volatilities.

Element: A substance that cannot be separated into simpler substances by chemical means; e.g.: C, N, Fe, Na, etc.

Notation of E.: Elements of the periodic table are assigned with a mass- and atomic number to quantify its number of protons (**Z**) and number of protons and neutrons (**A**); see chemistry atom.

Representative E.: Elements in groups 1A through 7A, all of which have incompletely filled s or p subshell of highest principal quantum number.

Experiment: A test carried out under carefully controlled conditions.

Group: The elements in a vertical column of the periodic table - see period.

Isomere: (Gk. isos, equal; meros, part) One of a group of compounds identical in atomic composition but differing in structural arrangement; e.g., glucose and fructose

Matter: Anything that occupies space and possesses mass.

Properties of M.: 1) A physical property can be measured and observed without changing the composition or identity of a substance. 2) in order to observe it, a chemical change has to be carried out; e.g.: hydrogen burns with oxygen to form water - see chemistry-solid -liquid, -gas.

- **Extensive P.o.M.:** A physical property of a substance that depends of the size of the sample; e.g.: mass, internal energy, enthalpy, entropy, etc.

- **Intensive P.o.M.:** A physical property of a substance that is dependent of the size of the sample; e.g.: density, molar volume, temperature, etc.

Mineral: A naturally occurring substance with a range of chemical composition.

Mixture: A type of matter that consists of more than one substance and may be separated into components by making use of the different physical properties.

Heterogeneous M.: A mixture in which the individual components, although mixed together, lie in distinct regions, even on a microscopic scale; e.g.: a mixture of sand and sugar, etc.

Homogenous M.: A mixture in which the individual components are uniformly mixed, even on an atomic scale; e.g.: air, solutions, etc.

Ore: The natural mineral source of a metal; e.g.: Fe₂O₃, hematite; etc.

Period: The horizontal row of the periodic table; the number of the period is equal to the principal quantum number of the valence shell of the atoms; periods from 1 to 7.

P. Table: A chart in which the elements are arranged in order of atomic number and divided into groups and periods in a manner that shows the relationships between the properties of the elements.

Group: The vertical column of the P.T.; the number of the group is equal to the number of electrons in the valence shell of the atoms; e.g.: groups from I to VIII.

Groupings within the P.T.:

Actinid Series: Elements that have incompletely filled 5f subshells or readily give rise to cations that have incompletely filled 5f subshells.

Alkali Metal: The group 1A elements like Li, Na, K, Rb, Cs, and Fr.

Alkali Earth Metal: A group of grayish-white, malleable metals easily oxidized in air, comprising Be, Mg, Ca, Sr, Ba, and Ra.

Halogen: The nonmetallic elements in Group 7A, F, Cl, Br, I, and At.

Lanthanide Series (are earth series): Elements that have incompletely filled 4f subshells or readily give rise to cations that have incompletely filled 4f subshells.

Metalloid: An element with properties intermediate between those of metals and nonmetals; i.e.: it has the physical appearance and properties of a metal but behaves chemically like a nonmetal; e.g.: arsenic, polonium, etc.

Metal: 1) A substance that conducts electricity as well as heat, has a metallic luster, is malleable and ductile, forms cations and has basic oxides. 2) A metal consists of cations held together by a sea of electrons (have the tendency to form positive ions in ionic compounds); i.e.: iron, copper, uranium, etc.

Noble Gas: Nonmetallic elements in group 8A; He, Ne, Ar, Kr, Xe, and Rn.

Transition Metal.: Elements that have incompletely filled d subshells or readily give rise to cations that have incompletely filled d subshells; i.e.: it belongs to the central part of the periodic table, between Groups II and III; a member of the d-block of the periodic table; i.e.: vanadium, iron, gold, etc.

ppm - **Parts per Million:** 1) The ratio of the mass of a solute to the mass of the solvent, multiplied by 10^6 ; 2) The mass percentage composition multiplied by 10^4 ;

Phase: A particular state of matter; a homogeneous part of a system in contact with other parts of the system but separated from them by a well defined boundary; a substance may exist in solid, liquid, and gas phases and in certain cases, in more than one solid phase; e.g.: white and gray tin are two solid phases of tin; ice, liquid, and vapor are three phases of water; etc.

P. Change: Transformation from one phase to another.

P. Diagram: A map divided into regions that tells us which phase is the most stable under corresponding conditions of temperature and pressure. The lines separating the areas in the diagram are called phase boundaries. The points on a phase boundary show the conditions under which two phases coexist in dynamic equilibrium.

Distinct Points in a P.D.:

K_p , **Critical point:** The liquid-vapor boundary terminates at this point; for water at $101,3[\text{kPa}] = 100[^\circ\text{C}]$.

T_p , **Triple point:** The point where 3 phase boundaries meet; for water it occurs at $0,61[\text{kPa}]$ and $0,01[^\circ\text{C}]$; under these conditions all three phases (ice, water, vapor) coexist in dynamic equilibrium.

B_p , **Boiling point:** The temperature at which a liquid boils when the atmospheric pressure is $1[\text{atm}]$ or $101,3[\text{kPa}]$.

F_p , **Freezing point:** The temperature at which it freezes (melts) at $1[\text{atm}]$; i.e.: liquid and solid state are in dynamic equilibrium.

Significant Figures: The number of meaningful digits in a measured or calculated quantity.

STP - **Standard Temperature and Pressure:** $0[^\circ\text{C}] = 273,15[\text{K}]$ and $1[\text{atm}] = 101,325[\text{kPa}]$.

Substance: A form of matter that has a definite or constant composition (the number and type of basic units present) and distinct properties.

SI - Base Units: (F. système international): Modern system of definitions and metric notation, now spreading throughout the academic, industrial, and commercial community; these are: ampere, area, joule, kelvin, kilogram, meter, mole, newton, rad, second, volume.

- n **Amount of Substance** [mol]
The amount of substance that contains as many elementary entities (atoms, molecules, or other particles) as there are atoms in exactly 12 grams of ^{12}C isotope.
- I **Electric Current** [ampere, = A]
the flow of 1 coulomb ($1\text{C} = 6.25 \times 10^{18}$ electrons) of charge/s.
- F **Force** [$\text{kg} \cdot \text{m}/\text{s}^2$] [newton, = N]
the force that will give an object of 1 kg an acceleration of $1 \text{ m}/\text{s}^2 = [\text{kg m}/\text{s}^2]$.
- l **Length** [meter, = m]
the length of the path traveled by light in vacuum during a time of $1/299792458$ of a second. Area: [m^2].
Volume: [m^3] Quantity of space an object occupies.
- L **Light Intensity** [$\text{N} \cdot \text{m}/(\text{s} \cdot \text{sr})$] = [$\text{J}/(\text{s} \cdot \text{sr})$] = [W/sr] = [candela, Cd]
light intensity of a monochromatic radiation with a frequency of 540×10^{12} oscillations /s [Hz] with a power in the direction equal to $1/683$ [Js or W/steradian].
- m,M **Mass** [kilogram, = kg]
one kilogram is the amount of mass in 1 liter of water at 4°C .
- rad **Radian:** The radian is the 2D plane angle between two radii of a circle which cut off on the circumference an arc equal in length to the radius: $1 \text{ rad} = 57.3^\circ$; $\pi \text{ rad} = 180^\circ$;
sr - steradian: Is the solid 3D angle which, having its vertex in the center of a sphere, cuts off an area equal to that of a flat square with sides of length equal to the radius of the sphere.
- T **Thermodynamic Temperature** [kelvin, = K]
defined to be $1/273.15$ the thermodynamic temperature of the triple point of water; ice melts therefore at 273.15 K and water boils at $373,15 \text{ K}$ (both at atmospheric pressure).
- t **Time:** Second [s]; the time taken by a ^{133}Cs -atom to make 9 192 631 770 vibrations.
- W **Work** [$\text{N} \cdot \text{m}$] [joule, = J]
the specific heat of work at 15°C is given as $4185.5 \text{ J}/\text{kg} \cdot \text{C}$ done by a force of 1 newton acting over a distance of 1 meter.

SI-Derived Units:

α	dissociation constant of acid / base	[-]	
β	mass concentration (solute over solvent)	[g/l]	
ρ	density	[kg/m ³]	[g/l]
η	coefficient of efficiency	[-]	
λ	wavelength	[m]	
π	circle's constant	= 3.14159	[-]
π	osmotic pressure	[Pa]	
τ	level of titration	[-]	
μ	dipole moment [C·m]	[D]	
\mathfrak{R}, R	gas constant	= 8.314 510	[J/(mol·K)]
A	area, cross-sectional area	[m ²]	
b	molality (solute over solvent)	[mol/g]	
c	speed of light	2.99792458·10 ⁸ \approx 3·10 ⁸	[m/s]
c	specific heat capacity	[N·m/(kg·K)]	[J/(kg·K)]
c	molar concentration	[mol/l]	
d	distance	[m]	
D	debye	= 3.336·10 ⁻³⁰	[C·m]
e	charge of an electron	= 1.602 177 3349·10 ⁻¹⁹	[C]
E	energy [kg·m ² /s ²] = [N·m] = [Pa·m ³]	[joule, J]	1eV = 1.6022·10 ⁻¹⁹ [J]
F	Faraday's constant	= 96485	[C/mole] 1[cal] = 4.178·10 ³ [J]
h	plank's constant	= 6.626 075 540·10 ⁻³⁴	[J·s]
H	ΔH enthalpy	[J/mol]	
k	Henry's law variable	[mol/(l·Pa)]	
k	kinetic rate law	[l/(mol·s)] or [1/s] or [mol/(l·s)]	varies, depending upon the rate order.
k _B	Boltzman's constant	= 1,380 658 12·10 ⁻²³	[J/K]
k _F	Coulomb's force constant	= 8.987 551 79·10 ⁹	[N·m ² /C ²] \approx 9x10 ⁹ = 1/(4· π · ϵ_0)
K _{A,B}	dissociation constant of acids and bases	[mol/l]*	
K _C	equilibrium constant for molar concentrations	[mol/l]*	
K _f	complex formation constant	[mol/l]*	
K _P	equilibrium constant for partial pressures	[Pa/l]*	
K _{sp}	solubility product	[mol/l]*	
K _w	water auto-dissociation constant	= 1·10 ⁻¹⁴	[mol ² /l ²]
l	length	[m]	
m _e	mass of an electron	= 9.109 389 754·10 ⁻³¹	[kg]
m _p	mass of a proton	= 1.672 623 110·10 ⁻²⁷	[kg]
m _n	mass of a neutron	= 1.674 928 610·10 ⁻²⁷	[kg]
M, M _r	molecular mass	[g/mol], [amu]	
n	molar amount	[mol]	
N _A	Avogadro's constant	= 6,022 136 736·10 ²³	[1/mol]
p	impulse	[kg·m/s]	
p	pressure [N/m ²] 1 bar = 10 ⁵ [Pa]	[pascal, Pa] 1atm = 1.01325[bar] = 101.325[kPa]	
q	electric charge of an electron (see e)	[A·s]	[coulomb, C]
Q	heat capacity	[J]	
Q _C	reaction quotient	[mol/l]*	
r	radius	[m]	
s	specific heat capacity	[J/(kg·K)]	
S	entropy [kg·m/(s ² ·K)] = [N·m/K]	[J/K]	
U	internal heat	[J]	
v	velocity	[m/s]	
v _(Ax)	kinetic reaction rate	[mol/(l·s)]	
V	electric potential [J/C]	[volt, V]	
V	volume	[m ³]	
w	percent by mass	[%]	
x	molar fraction	[-]	

SI-Prefixes: 10¹⁸ *exa-* E; 10¹⁵ *peta-* P; 10¹² *tera-* T; 10⁹ *giga-* G; 10⁶ *mega-* M; 10³ *kilo-* k; 10 *deka-* d; 10⁻³ *milli-* m; 10⁻⁶ *micro-* μ ; 10⁻⁹ *nano-* n; 10⁻¹² *pico-* p; 10⁻¹⁵ *femto-* f; 10⁻¹⁸ *atto-* a;

*) varies, depending upon stoichiometric coefficient of products and reactants; e.g.: [l/mol], [mol²/l²], etc.

Periodic Table of Elements

Group	I		II												III	IV	V	VI	VII	VII	I
Period																					
1	H																	He			
2	Li	Be													B	C	N	O	F	Ne	
3	Na	Mg											Al	Si	P	S	Cl	Ar			
4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr			
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe			
6	Cs	Ba	La	*)	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn		
7	Fr	Ra	Ac	**)																	
Block	s		d	f	d										p						

*)	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
**)	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

Element	amu	Z	A	Element	amu	Z	A	Element	amu	Z	A
Ac	Actinium	(227)	89	H	Hydrogen	1.008	1	Pm	Promethium	(147)	61
Ag	Silver	107.9	47	He	Helium	4.003	2	Po	Polonium	(209)	84
Al	Aluminium	26.98	13	Hf	Hafnium	178.5	72	Pr	Praseodymium	140.9	59
Am	Americum	(243)	95	Hg	Mercury	200.6	80	Pt	Platinum	195.1	78
Ar	Argon	39.95	18	Ho	Holmium	165.0	67	Pu	Plutonium	(209)	94
As	Arsenic	74.92	33	Ga	Gallium	69.72	31	Ra	Radium	226.0	88
At	Astatine	(210)	85	Gd	Gadolinium	157.3	64	Rb	Rubidium	85.47	37
Au	Gold	197.0	79	Ge	Germanium	72.59	32	Re	Rhenium	186.2	75
B	Boron	10.81	5	I	Iodine	126.9	53	Rh	Rhodium	102.9	45
Ba	Barium	137.3	56	In	Indium	114.8	49	Rn	Radon	(222)	86
Be	Beryllium	9.012	4	Ir	Iridium	192.2	77	Ru	Ruthenium	101.1	44
Bi	Bismuth	209.0	83	K	Potassium	39.10	19	S	Sulfur	32.07	16
Bk	Berkilium	(247)	97	Kr	Krypton	83.80	36	Sc	Scandium	44.96	21
Bo	Boron	10.81	5	La	Lanthanum	138.9	57	Se	Selenium	78.96	34
Br	Bromine	79.90	35	Li	Lithium	6.941	3	Si	Silicon	28.09	14
C	Carbon	12.01	6	Lr	Lawrencium	(260)	103	Sb	Antimony	121.8	51
Ca	Calcium	40.08	20	Lu	Lutetium	175.0	71	Sm	Samarium	150.4	62
Cd	Cadmium	112.4	48	Md	Mendelevium	(258)	101	Sn	Tin	118.7	50
Ce	Cerium	140.1	58	Mg	Magnesium	24.30	12	Sr	Strontium	87.62	38
Cf	Californium	(249)	98	Mn	Manganese	54.94	25	Ta	Tantalum	180.9	73
Cl	Chlorine	34.45	17	Mo	Molybdenum	95.94	42	Tb	Terbium	158.9	65
Cm	Curium	(247)	96	N	Nitrogen	14.01	7	Tc	Technetium	(98)	43
Cs	Cesium	132.9	55	Na	Sodium	22.99	11	Te	Tellurium	127.6	52
Cr	Chromium	52.00	34	Nb	Noibium	92.91	41	Th	Thorium	232.0	90
Co	Cobalt	58.93	27	Ne	Neon	20.18	10	Ti	Titanium	47.88	22
Cu	Copper	63.55	29	Nd	Neodymium	144.2	60	Tl	Thallium	168.9	69
Dy	Dysprosium	162.5	66	Ni	Nickel	58.69	28	Tm	Thulium	168.9	69
Er	Erbium	167.3	68	No	Nobelium	(259)	102	U	Uranium	238.0	92
Es	Einsteinium	(252)	99	Np	Neptunium	237.0	93	V	Vanadium	50.94	23
Eu	Europium	152.0	63	O	Oxygen	16.00	8	W	Tungsten	183.8	75
F	Flourine	19.00	9	Os	Osmium	190.2	76	Xe	Xenon	131.3	54
Fe	Iron	55.85	26	P	Phosphorus	30.97	15	Zn	Zinc	65.39	30
Fm	Fermium	(257)	100	Pa	Protactinium	231.0	91	Zr	Zirconium	91.22	40
Fr	Francium	(223)	87	Pb	Lead	207.2	82	Y	Yttrium	88.91	39
				Pd	Palladium	106.4	46	Yb	Ytterbium	173.0	70

Glossary - Chemistry (Atom)

Atom: The basic unit of an element that can enter into chemical combination (see physics - nuclear).

Z - Atomic Number: The number of protons in the nucleus of an atom; it determines the identity of the element and the number of electron in the neutral atom.

A. Orbital: see orbital.

A. Radius: $\frac{1}{2}$ the distance between the two nuclei in two adjacent atoms of the same element in a metal. For elements that exist as diatomic units, the atomic radius is $\frac{1}{2}$ the distance between the nuclei of the two atoms in a particular molecule; The A.R. decreases from left to right across the period of the periodic table, whereas it increases going down a group of the periodic table.

A. Weight: The weight of a representative atom of an element relative to the weight of an atom of ^{12}C , which has been assigned the value 12 - see stoichiometry - mass.

Average Atomic Mass: The average mass of the naturally occurring mixture of isotopes as listed in the periodic table of elements - see chemistry introduction - periodic table.

Parts of the atom:

Electron: A subatomic particle that has a very low mass and carries a single negative electric charge (see physics - introduction). $m_e = 9.1095 \cdot 10^{-28}$ [g]
 $q_e = -1.6022 \cdot 10^{-19}$ [C]

Neutron: A subatomic particle that bears no electric charge. Its mass is slightly greater than a proton's. $m_n = 1.67495 \cdot 10^{-24}$ [g]

Proton: A subatomic particle having a single positive electric charge. The mass of a proton is about 1840 times that of an electron. $m_p = 1.67252 \cdot 10^{-24}$ [g]
 $q_p = +1.6022 \cdot 10^{-19}$ [C]

Aufbau Principle: As protons are added one by one to the nucleus to build up the elements, electrons similarly are added to the atomic orbitals (see table below).

Bohr's Atomic Model: Energy of the electron orbiting the hydrogen nucleus is quantized into fixed values; i.e. electron can only „move“ along fixed steps of energy (no values in-between)

$$E = R \cdot h \cdot c / n^2 \quad [N \cdot m] = [J]$$

R, Rydberg c. = $2.18 \cdot 10^{-18}$ [J]
n, principal quantum number
c, speed of light $3 \cdot 10^8$ [m/s]
h, Planck c. = $6.63 \cdot 10^{-34}$ [J·s]

Dalton's Atomic Theory: Hypotheses about the nature of matter.

- Elements are composed of atoms. All atoms of a given element are identical (except for isotopes), having the same size, mass, and chemical properties. Atoms of one element are different from atoms of all other elements.
- Compounds are composed of atoms of more than one element. In any compound, the ratio of numbers of atoms is either an integer or a simple fraction.
- Chemical reactions involve only the separation, combination, or rearrangement of atoms; it does not result in their creation or destruction (except nuclear reaction).

De-Broglie's Hypothesis: The proposal that every particle has wavelike properties; hence it is impossible to determine both position and velocity simultaneously:
 $\lambda = h / (m \cdot c) = h / p$ [m]

h, Planck c. = $6.63 \cdot 10^{-34}$ [J·s]
m, mass [kg]
c, speed of light = $3 \cdot 10^8$ [m/s]
p, impulse [kg·m/s]

Electron: see atom - parts of.

E. Affinity: The energy change when an electron is accepted by an atom (or ion) in the gaseous state.

E. Configuration: The distribution of electrons among the various orbitals in an atom or molecule.

E. Density: The probability that an electron will be found at a particular region in an atomic orbital.

E. per Shell: The principles outlined by both the atomic Bohr model and the rules of the quantum number, only a certain amount of e can occupy each orbital:

$$e_N = 2 \cdot n^2$$

e_N , number of e/shell [-]
n, principal shell number [-]

Excited State: A state that has higher energy than the ground state.

Exclusion Principle: see pauli.

FC - Formal Charge: 1) The electric charge of an atom assigned on the assumption that there is perfect covalent bonding. 2) FC = number of valence electrons in the free atom - number of lone-pair electrons - $\frac{1}{2}$ x number of shared electrons; see chemistry molecule.

Ground State: The lowest energy state of a system.

Heisenberg uncertainty Principle: It is impossible to know simultaneously both momentum and the position of a particle with certainty (see physics - nuclear).
 $\Delta x \cdot \Delta p = h / (4 \cdot \pi) = \Delta x \cdot \Delta(m \cdot c) \cdot h / (4 \cdot \pi)$ [kg·m²/s]

x, position [m]
m, mass [kg]
c, speed of light = $3 \cdot 10^8$ [m/s]
h, Planck c. = $6.63 \cdot 10^{-34}$ [J·s]

Hund's Rule: The most stable arrangement of electrons in subshells is the one with greatest number of parallel spins; e.g.: C with 6 electrons, where the $2p^2$ of the $1s^2 2s^2$ are parallel - compare chemistry molecule - hybridization.

Inert Pair Effect: The two relatively stable and unreactive outer s electrons.

Ion: An atom or molecule that has lost or gained one or more electrons, and thus becomes positively or negatively charged; i.e.: Al^{3+} (mono-atomic ion), SO_4^- (poly-atomic ion); see also chemistry-molecule & thermochem.

Anion: An ion with a net negative charge; characteristic of non-metallic elements, e.g.: F^- , SO_4^- , etc.

Cation: An ion with a net positive charge; characteristic of metallic elements, e.g.: Na^+ , NH_4^+ , Al^{3+} , etc.

Ionic Radius: The radius of a cation or an anion as measured in an ionic compound (see table below); it decreases from left to right across the a period and increases down a group of the periodic table; in the same period, anions generally have larger radii than cations.

Ionization: Conversion to cations by the removal of electrons; see chemistry-thermochemistry ΔH .

e.g.: $K(g) \rightarrow K^+(g) + e^-(g)$

Isoelectronic Species: Species with the same number of atoms and the same number of valence electrons;

e.g.: SO_2 and O_3 ; Cl^- and Ar ; Na^+ and Ne , NO_3^- and CO_3^{2-} (Ar , Ne are noble gases).

Isotope: One of several possible forms of a chemical element that differ from others in the number of neutrons in the atomic nucleus, but not in chemical properties e.g.: ^{21}Na , ^{22}Na , ^{23}Na .

I. Dating: The determination of the ages of objects by measuring the activity of the radioactive isotopes they contain; particularly, radiocarbon dating by using ^{14}C .

Magnetism: A materials ability to store magnetic energy after being magnetized describes 3 distinct classes:

Diamag.: Repelled by a magnet; a diamagnetic substance contains only paired electrons, resulting in a magnetic flux which is lower inside than outside ($\mu < 1$, slightly repelling when exposed to a magnetic field; temperature independent) - H_2 , Cu , Hg ...

Ferromag.: Strongly attracted by a magnet; the unpaired spins in a ferromagnetic substance are aligned in a common direction; the magnetic flux inside is extremely high compared to outside; ($\mu \gg 1$, the higher the temperature the less ferromagnetic will be the substance) - Fe , Ni , Co ...

Paramagnet.: Attracted by a magnet. A paramagnetic substance contains one or more unpaired electrons; resulting in a magnetic flux which is higher inside than outside ($\mu \geq 1$, slightly attracting, the lower the temperature the more ferromagnetic the substance will become) - O_2 , Cr , Pt ...

Notation of Elements: Elements of the periodic table are assigned with a mass- and atomic number to quantify its number of protons (**Z**) and number of protons and neutrons (**A**).

Orbital: The idealized representation of electron moving around its nucleus.

Atomic O.: The wave function (ψ) of an electron in an atom (see quantum number).

Degenerate O.: Orbitals that have the same energy.

Electron Configuration: The occupancy of orbitals in an atom or molecule; i.e.: $N 1s^2 2s^2 2p^3$, $N_2 \sigma^2 \sigma^{*2} \pi^4 \sigma^2$ (valence electrons only); see chemistry molecule.

Partial Charge: A charge arising from small shifts in the distribution of electrons.

Pauli Exclusion Principle: No two electrons in an atom can have the same four quantum numbers.

Photon: A particle of light (see physics-optics); The energy contained in a photon can be put as follows:

$$E = h \cdot f = h \cdot c / \lambda \quad [J]$$

Quantum: The smallest quantity of energy that can be emitted,

or absorbed in the form of electromagnetic radiation

(see physics-optics).

$$c = \lambda \cdot f \quad [m/s]$$

h , Planck c. $6.63 \cdot 10^{-34}$ [J·s]

c , speed of light $3 \cdot 10^8$ [m/s]

λ , wavelength [m]

f , frequency [1/s] [Hz]

Quantum Number: The quantum numbers of quantum mechanics which describe the distribution of electrons, labels the state of the electron and specifies the value of a property in atoms.

Paired Electrons: Two electrons with opposed spins ($\uparrow\downarrow$).

Parallel Electrons: Electrons with spins in the same direction ($\uparrow\uparrow$).

1. **n - Principal QN.**(shell number): The average distance of the electron from the nucleus in a particular orbital; can have integral values of 1, 2, 3, and so forth (higher values \approx greater average distance) e.g.: 1 = 1st period, 7 = 7th period;

2. **l - Angular Momentum QN.**(subshell of one shell): Its value reflects the orbital shape; it correlates with n ; ($l = n - 1$); which reveals 0 for the **s**-, 1 for **p**-, 2 for **d**- 3 for **f**-, 4 for **g**- 5 for the **h**-orbital.

3. **m_l - Magnetic QN.:** Describes the orientation of the orbital in space; depends upon the value of l ; ($m_l = 2 \cdot l + 1$), e.g.: $m_l = 1$ a sphere; $m_l = 3$ gives $-1/0/+1$ (x,y,z-orientation); $m_l = 5$ gives $-2/-1/0/1/2$ etc.

4. **m_s - Electron Spin QN.:** According to the electromagnetic theory, spinning electrons possess a magnetic orientation; ($m_s = n$) m_s can either be $-1/2$ (\downarrow) or $+1/2$ (\uparrow); with $m_s = 3$ giving 3 magnetic spins: $-1/2 / +1/2 / -1/2$.

$\Delta\psi$ - **Schrödinger Wave Equation**: The fundamental equation of quantum mechanics; it interprets the wave nature of material particles in terms of probability wave amplitudes.

It specifies the possible energy state the electron can occupy and identifies the corresponding wave function (ψ).

These energy states and wave functions are characterized by the quantum numbers; The concept of electron density gives the probability that an electron will be found in a particular region of an atom:

$$\Delta\psi + (8 \cdot \pi^2 \cdot m / h^2) \cdot (E - PE) \cdot \psi = 0$$

Series: A family of spectral lines arising from transitions that have one state in common (see spectrum):

Bracket S.: Energy spectrum located in the infrared ($n_i = 4$)

Paschen S.: Energy spectrum located in the infrared ($n_i = 3$)

Balmer S.: A family of spectral lines (some of which lie in the visible region, others in the UV band), in the spectrum of atomic hydrogen; ($n_i = 2$)

Lyman S.: Energy spectrum located entirely in the UV ($n_i = 1$)

Shell: All the orbitals of a given principal quantum number $>n<$; i.e.: the single 2s and three 2p orbitals of the shell with $n = 2$;

S.-Nomenclature: s - sharp; p - principal; d - diffuse; f - fundamental

Subshell: All the atomic orbitals of a given shell of an atom with the same value of the quantum number $>l<$; i.e.: the five 3d orbitals of an atom.

Spectroscope: An optical device that separates light into its constituent frequencies in form of spectral lines.

Spectrum: 1) The frequencies or wavelengths of electromagnetic radiation emitted or absorbed by substances
2) The splitting of white light into its distinct components.

Emission S. of Hydrogen: A line spectrum of radiation emitted by the excited hydrogen electron. (integers of $n_i > n_f$) - see series

$$\Delta E = R \cdot h \cdot c (1/n_i^2 - 1/n_f^2) \quad [J]$$

S. Line: Radiation of a single wavelength emitted or absorbed by an atom or molecule.

Line Spectra: Spectra produced when radiation is absorbed or emitted by substances only at some wavelengths.

Spin: The intrinsic angular momentum of an electron; the spin cannot be eliminated and may occur in only two senses, denoted as \uparrow and \downarrow - see quantum number.

$\Delta\psi$, Laplace operator $[-1/m]$

ψ , wave function $[m]$

ψ^2 , probability of density $[-]$

PE, potential energy $[J]$

E, total energy $[J]$

m, mass $[kg]$

h, Planck c. = $6.63 \cdot 10^{-34}$ $[J \cdot s]$

π , circle's constant = 3.142 $[-]$

R, Rydberg constant $2.18 \cdot 10^{-18}$ $[J]$

n_i , initial principal QN

n_f , final principal QN

h, Planck's c. = $6.63 \cdot 10^{-34}$ $[J \cdot s]$

c, speed of light = $3 \cdot 10^8$ $[m/s]$

Atomic and ionic radii of common elements before and after ionization [pm]:

	I	II			III	IV	V	VI	VII	VIII	
1	H									He (30) He (154)	
2	Li (142) Li ⁺ (60)	Be (112) Be ⁺ (31)			B (88)	C (77)	N (70)	O (66) O ²⁻ (140)	F (64) F (138)	Ne	
3	Na (186) Na ⁺ (95)	Mg (160) Mg ²⁺ (65)			Al (143) Al ³⁺ (50)	Si (117)	P (110)	S (104) S ²⁻ (184)	Cl (99) Cl (181)	Ar	
4	K (231) K ⁺ (133)	Ca (197) Ca ²⁺ (97)	Sc (160) Sc ³⁺ (81)	Cu (128) Cu ⁺ (96)	Zn (133) Zn ²⁺ (74)	Ga (122) Ga ³⁺ (62)	Ge (122) Ge ⁴⁺ (53)	As (121)	Se (117) Se ²⁻ (198)	Br (114) Br (195)	Kr
5	Rb (244) Rb ⁺ (148)	Sr (215) Sr ²⁺ (113)	Y (180) Y ³⁺ (95)	Ag (144) Ag ⁺ (126)	Cd (149) Cd ²⁺ (97)	In (162) In ³⁺ (81)	Sn (140) Sn ⁴⁺ (71)	Sb (141)	Te (137) Te ²⁻ (221)	I (133) I (216)	Xe
6	Cs (262) Cs ⁺ (169)	Ba (217) Ba ²⁺ (135)	La (188) La ³⁺ (115)	Au (144) Au ⁺ (137)	Hg (150) Hg ²⁺ (110)	Tl (171) Tl ³⁺ (95)	Pb (175) Pb ⁴⁺ (84)	Bi (146)	Po (140)	At (140)	Rn

The order in which atomic orbitals are filled according to the Aufbau principle:

Core	I	II											III	IV	V	VI	VII	VIII			
	1s	→																			
	2s		→											→	2p						
	3s		→											→	3p						
	4s		3d	→	→	3d									4p						
	5s		4d	→	→	d4									5p						
	6s		5d	4f											6p						
	7s		6d	5f																	
Rn	2 s-b g.		1	14 f-block grp.						(1+) 9 d-block groups						6 p-block groups					

The ground-state electron configuration of the elements:

Element	Electron config.	Element	Electron config.	Element	Electron config.
Ac	Actinium [Rn] 7s ² 6d ¹	H	Hydrogen [] 1s ¹	Pm	Promethium [Xe] 6s ² 4f ⁵
Ag	Silver [Kr] 5s ¹ 4d ¹⁰	He	Helium [] 1s ²	Po	Polonium [Xe] 6s ² 4f ¹⁴ 5d ¹⁰ 6p ⁴
Al	Aluminium [Ne] 3s ² 3p ¹	Hf	Hafnium [Xe] 6s ² 4f ¹⁴ 5d ²	Pr	Praseodymium [Xe] 6s ² 4f ³
Am	Americum [Rn] 7s ² 5f ⁷	Hg	Mercury [Xe] 6s ² 4f ¹⁴ 5d ¹⁰	Pt	Platinum [Xe] 6s ¹ 4f ¹⁴ 5d ⁹
Ar	Argon [Ne] 3s ² 3p ⁶	Ho	Holmium [Xe] 6s ² 4f ¹¹	Pu	Plutonium [Rn] 7s ² 5f ⁶
As	Arsenic [Ar] 4s ² 3d ¹ 4p ³	Ga	Gallium [Ar] 4s ² 3d ¹ 4p ¹	Ra	Radium [Rn] 7s ²
At	Astatine [Xe] 6s ² 4f ¹⁴ 5d ¹⁰ 6p ⁵	Gd	Gadolinium [Xe] 6s ² 4f ⁷ 5d ¹	Rb	Rubidium [Kr] 5s ¹
Au	Gold [Xe] 6s ¹ 4f ¹⁴ 5d ¹⁰	Ge	Germanium [Ar] 4s ² 3d ¹ 4p ²	Re	Rhenium [Xe] 6s ² 4f ¹⁴ 5d ⁵
B	Boron [He] 2s ² 2p ¹	I	Iodine [Kr] 5s ² 4d ¹⁰ 5p ⁵	Rh	Rhodium [Kr] 5s ¹ 4d ⁸
Ba	Barium [Xe] 6s ²	In	Indium [Kr] 5s ² 4d ¹⁰ 5p ¹	Rn	Radon [Xe] 6s ² 4f ¹⁴ 5d ¹⁰ 6p ⁶
Be	Beryllium [He] 2s ²	Ir	Iridium [Xe] 6s ² 4f ¹⁴ 5d ⁷	Ru	Ruthenium [Kr] 5s ¹ 4d ⁷
Bi	Bismuth [Xe] 6s ² 4f ¹⁴ 5d ¹⁰ 6p ³	K	Potassium [Ar] 4s ¹	S	Sulfur [Ne] 3s ² 3p ⁴
Bk	Berkilium [Rn] 7s ² 5f ⁹	Kr	Krypton [Ar] 4s ² 3d ¹ 4p ⁶	Se	Selenium [Ar] 4s ² 3d ¹ 4p ⁴
Br	Bromine [Ar] 4s ² 3d ¹ 4p ⁵	La	Lanthanum [Xe] 6s ² 5d ¹	Si	Silicon [Ne] 3s ² 3p ²
C	Carbon [He] 2s ² 2p ²	Li	Lithium [He] 2s ¹	Sb	Antimony [Kr] 5s ² 4d ¹⁰ 5p ³
Ca	Calcium [Ar] 4s ²	Lr	Lawrencium [Rn] 7s ² 5f ¹⁴ 6d ¹	Sc	Scandium [Ar] 4s ² 3d ¹
Cd	Cadmium [Kr] 5s ² 4d ¹⁰	Lu	Lutetium [Xe] 6s ² 4f ¹⁴ 5d ¹	Sm	Samarium [Xe] 6s ² 4f ⁶
Ce	Cerium [Xe] 6s ² 4f ¹ 5d ¹	Md	Mendelevium [Rn] 7s ² 5f ¹³	Sn	Tin [Kr] 5s ² 4d ¹⁰ 5p ²
Cf	Californium [Rn] 7s ² 5f ¹⁰	Mg	Magnesium [Ne] 3s ²	Sr	Strontium [Kr] 5s ²
Cl	Chlorine [Ne] 3s ² 3p ⁵	Mn	Manganese [Ar] 4s ² 3d ⁵	Ta	Tantalum [Xe] 6s ² 4f ¹⁴ 5d ³
Cm	Curium [Rn] 7s ² 5f ⁷ 6d ¹	Mo	Molybdenum [Kr] 5s ¹ 4d ⁵	Tb	Terbium [Xe] 6s ² 4f ⁹
Cs	Cesium [Xe] 6s ¹	N	Nitrogen [He] 2s ² 2p ³	Tc	Technetium [Kr] 5s ² 4d ⁵
Cr	Chromium [Ar] 4s ¹ 3d ⁵	Na	Sodium [Ne] 3s ¹	Te	Tellurium [Kr] 5s ² 4d ¹⁰ 5p ⁴
Co	Cobalt [Ar] 4s ² 3d ⁷	Nb	Niobium [Kr] 5s ¹ 4d ⁴	Th	Thorium [Rn] 7s ² 6d ²
Cu	Copper [Ar] 4s ¹ 3d ¹⁰	Ne	Neon [He] 2s ² 2p ⁶	Ti	Titanium [Ar] 4s ² 3d ²
Dy	Dysprosium [Xe] 6s ² 4f ¹⁰	Nd	Neodymium [Xe] 6s ² 4f ⁴	Tl	Thallium [Xe] 6s ² 4f ¹⁴ 5d ¹⁰ 6p ¹
Er	Erbium [Xe] 6s ² 4f ¹²	Ni	Nickel [Ar] 4s ² 3d ⁸	Tm	Thulium [Xe] 6s ² 4f ¹³
Es	Einsteinium [Rn] 7s ² 5f ¹¹	No	Nobelium [Rn] 7s ² 5f ¹⁴	U	Uranium [Rn] 7s ² 5f ³ 6d ¹
Eu	Europium [Xe] 6s ² 4f ⁷	Np	Neptunium [Rn] 7s ² 5f ⁴ 6d ¹	V	Vanadium [Ar] 4s ² 3d ³
F	Flourine [He] 2s ² 2p ⁵	O	Oxygen [He] 2s ² 2p ⁴	W	Tungsten [Xe] 6s ² 4f ¹⁴ 5d ⁴
Fe	Iron [Ar] 4s ² 3d ⁶	Os	Osmium [Xe] 6s ² 4f ¹⁴ 5d ⁶	Xe	Xenon [Kr] 5s ² 4d ¹⁰ 5p ⁶
Fm	Fermium [Rn] 7s ² 5f ¹²	P	Phosphorus [Ne] 3s ² 3p ³	Zn	Zinc [Ar] 4s ² 3d ¹⁰
Fr	Francium [Rn] 7s ¹	Pa	Protactinium [Rn] 7s ² 5f ² 6d ¹	Zr	Zirconium [Kr] 5s ² 4d ²
		Pb	Lead [Xe] 6s ² 4f ¹⁴ 5d ¹⁰ 6p ²	Y	Yttrium [Kr] 5s ² 4d ¹
		Pd	Palladium [Kr] 4d ¹⁰	Yb	Ytterbium [Xe] 6s ² 4f ¹⁴

the symbol [He] is called the helium-core and represents 1s²

[Ne]-core: 1s²2s²2p⁶

[Ar]-core: 1s²2s²2p⁶3s²3p⁶

[Kr]-core: 1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p⁶

[Xe]-core: 1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p⁶5s²4d¹⁰5p⁶

[Rn]-core: 1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p⁶5s²4d¹⁰5p⁶6s²4f¹⁴5d¹⁰6p⁶

Orbitals of the 1st four elements

period	Shell (n)	Sub-shell (l)	Orbital (m)	Subshell nomenclature	Number of orbitals per subshell
1	1	0	0	1s	1
2	2	0	0	2s	1
		1	-1, 0, +1	2p	3
3	3	0	0	3s	1
		1	-1, 0, +1	3p	3
		2	-2, -1, 0, +1, +2	3d	5
4	4	0	0	4s	1
		1	-1, 0, +1	4p	3
		2	-2, -1, 0, +1, +2	4d	5
		3	-3, -2, -1, 0, +1, +2, +3	4f	7

Electron-configuration of the 1st ten elements

Element	Orbital Diagram					Configuration
	1s	2s	2p _x	2p _y	2p _z	
₁ H	↑					1s ¹
₂ He	↑↓					1s ²
₃ Li	↑↓	↑				1s ² 2s ¹
₄ Be	↑↓	↑↓				1s ² 2s ²
₅ B	↑↓	↑↓	↑			1s ² 2s ² 2p ¹
₆ C	↑↓	↑↓	↑	↑		1s ² 2s ² 2p ²
₇ N	↑↓	↑↓	↑	↑	↑	1s ² 2s ² 2p ³
₈ O	↑↓	↑↓	↑↓	↑	↑	1s ² 2s ² 2p ⁴
₉ F	↑↓	↑↓	↑↓	↑↓	↑	1s ² 2s ² 2p ⁵
₁₀ Ne	↑↓	↑↓	↑↓	↑↓	↑↓	1s ² 2s ² 2p ⁶

Glossary - Chemistry (Molecule)

Bond: A link between bonds.

B. **Energy** (Bond Dissociation Energy): The enthalpy change required to break a bond in a mole of gaseous molecules - see thermochemistry.

B. **Length:** The distance between the centers of two bonded atoms in a molecule.

B. **Order:** see orbital-molecular.

Covalent B.: A bond in which two electrons are shared by two atoms. In a coordinate covalent bond the pair of electrons is supplied by one of the two bonded atoms;

e.g.: H-Cl, by sharing each others electrons, both obtain noble gas configuration.

- **π - Pi B.:** A covalent bond formed by sideways overlapping orbitals; its electron density is concentrated above and below the plane of the nuclei of the bonding atoms.
- **σ - Sigma B.:** A covalent bond formed by orbitals overlapping end-to-end; it has its electron density concentrated between the nuclei of the bonding atoms.
- **Coordinative (Dative) B.:** The lone pair of one compound is occupied by a cation;
e.g.: $\text{NH}_3 + \text{H}^+ \rightarrow \text{NH}_4^+$

Double B.: Two electron pairs shared by neighboring atoms; i.e.: one σ -bond and one π -bond.

Hydrogen B.: see forces.

Ionic B.: Electrostatic force that holds ions together in an ionic compound;

e.g.: $\text{Na} + \text{e}^- + \text{Cl} \leftrightarrow \text{Na}^+(\text{cation}) + \text{Cl}^-(\text{anion})$

Metallic B.: Ions (cations) in fixed in their proper lattice-position are held together by a sea of electrons (have the tendency to form positive ions in ionic compounds); atoms of metals give off electrons into the common pool of the conducting-band to obtain noble gas configuration.

Multiple B.: Bonds formed when two atoms share two or more pairs of electrons.

Nonbonding Electrons: Valence electrons that are not involved in covalent bond formation.

Peptide B.: A type of covalent bond joining amino acids in a polypeptide.

Polypeptide B.: Several amino acids linked together by peptide bonds.

Triple B.: Three electron pairs shared by two neighboring atoms; i.e.: one σ -bond and two π -bonds.

Catenation: The ability of atoms to form bonds with one another.

Compound: A substance composed of atoms of two or more elements chemically united in fixed properties.

Covalent Radius: The contribution of an atom to the length of a covalent bond - see table below;

e.g.: Cl-Cl ; $r = 198$ [pm]

q - **Charge:** Is a whole number multiple of one electron and cannot be quantized below it (1e or 1p-net charge); no net charge has ever been found - it can't be created or destroyed, only transformed (compare 1st law of thermodynamics) $e = 1.602 \cdot 18 \cdot 10^{-19}$ [As] = [C]. Like charges repel, opposite charges attract; electric charges can be isolated, magnet poles cannot.

FC - **Formal Charge:** 1) The electric charge of an atom assigned on the assumption that there is perfect covalent bonding. 2) FC = number of valence electrons in the free atom - number of lone-pair electrons - $\frac{1}{2}$ the number of shared electrons;

e.g.: overall e- availability of NO_3^- given by the octet rule would be: $e_v' = 4 \text{ elements} \cdot 8 = 32$;

N accounts for only 5, and O for (3·6) 24, plus one extra (e^-): $e_v'' =$ gives a total of 24;

the difference $\frac{1}{2} \cdot (32-24) = 8$ gives the maximum of possible bonds between these two atoms; in the case of NO_3^- $\frac{1}{2} \cdot 8 = 4$ bonds, with N in the middle (a planar Y-shaped molecular structure).

Partial C.: A charge arising from small shifts in the distribution of electrons, resulting in an electric dipole, see there.

Coulomb's Law: The potential energy between two ions is directly proportional to the product of their charges and inversly proportional to the distance of separation between them (see physics - electromagnetism).

The relationship among electrical force, charge, and distance:

$$F = k_F \cdot q_1 \cdot q_2 / d^2 \quad [\text{Nm}^2/\text{C}^2 \cdot \text{C}^2/\text{m}^2] = [\text{N}] \quad k_F, \text{coulomb c.} \approx 9 \cdot 10^9 \quad [\text{N} \cdot \text{m}^2/\text{C}^2]$$

If the charges are alike in sign, the force is repelling; $q, \text{charge} [\text{A} \cdot \text{s}] \quad 1.602 \cdot 10^{-19} \quad [\text{C}]$

If the charges are unlike, the force is attractive. $d, \text{distance} \quad [\text{m}]$

Cyanides: Compounds containing the CN^- ion.

Dipole: A molecule with a separate regions or net negative and net positive charge, so that one end acts as a positive pole and the other as a negative pole.

D. **Moment:** The electrostatic force required to align a dipolar molecule parallel to the electrostatic field; the force required increases as the separation of the molecular charges decreases; The product of charge and the distance between the charges in a molecule:

$$\mu = q \cdot d \quad [\text{C} \cdot \text{m}] = [\text{D}]; 1 \text{Debye} = 3.336 \cdot 10^{-30} [\text{C} \cdot \text{m}]$$

The overall DM is obtained by adding the individual vector-amounts of the involved atoms.

D. **Forces:** Forces that act between polar molecules; see force.

Electric D.: A positive charge next to an equal but opposite negative charge.

Induced D.: The separation of positive and negative charges in a neutral atom (or non-polar molecule) caused by the proximity of an ion or polar molecule.

Electronegativity: The measure of the ability of an atom to attract electron to itself when it is part of a compound; it is formed by the arithmetic sum of the actual bonding energy (A-B) and the hypothetical unpolar bond between A-B; the difference reveals the polar character of a bond; it increases from left to right across the period of the periodic table and decreases down a group; see table below.

Esters: Compounds that have the general formula R'COOR; R' can be H or an alkyl group or an aryl group.

Ether: An organic compound containing the R-O-R' linkage, where R and R' are alkyl and/or aryl groups.

F - **Force:** An influence that changes the state of motion of an object; i.e.: electrostatic force from an electric charge; a mechanical force from an impact, etc.;

$$F_L = m \cdot a \quad [\text{kg} \cdot \text{m/s}^2] = [\text{N}]$$

Intermolecular F. (Van der Waals): The close-ranging, relatively

weak attraction exhibited between atoms and molecules. Attractive forces between molecules that hold atoms together; the boiling point of an element or compound reflects the strength of the bonds involved:

- **Dipole-Dipole F.:** Forces acting between polar molecules; alike charges repel, opposite attract.

Hydrogen Bond.: A special, weak chemical bond (dipole-dipole interaction) between a hydrogen atom bonded to an atom of high electronegative elements (F, N, or O) and another atom of one of the three electronegative elements, accounting for the high boiling point of water HF, and NH₃.

- **Ion-Dipole F.:** The force of attraction between an ion and the opposite partial charge of the electric dipole of a polar molecule; e.g.: HCl, NaCl, AgCl, ect.

- **London F. (Dispersion):** The force of attraction that arises from the interaction b/w electric dipoles on neighboring polar / nonpolar molecules; it increases w/ the size of the molecule (molecular mass); e.g.: hydrocarbon with <4 central C-atoms, gaseous; >5-17<, liquid, >18 waxy solid.

Hybridization: The process of mixing the atomic orbitals in an atom (usually the central atom) to generate a set of new atomic orbitals - hybrid orbitals in a molecule (see orbital);

Hydrates: Compounds that have a specific number of water molecules attached to them; e.g.: CuSO₄·5H₂O;

Hydride: A binary compound of a metal or metalloid with hydrogen; the term is often extended to include all binary compounds of hydrogen.

Bonds w/ C,N,O: Generally those compounds wher a hydrogen atom is bonded to; hydrocarbons (C), ammonia (N), phoshine (P), water (O), hydrogen sulfide (S), hydrogen fluoride (F), etc.

Complex H.: Elements of the 3rd group linked to hydrogen in a tetrahedral bond; e.g.: NaBH₄, LiAlH₄.

Covalent H.: Hydrogen bonds are bonded covalently with metallic elements of group 3 as well with Be, and Mg; e.g.: BeH₂, GaH₃.

Metallic H.: A compound of certain d-block metals and hydrogen; in which H fills the gap in-between the metallic atoms; e.g.: TiH_x (x<2).

Molecular H.: A compound of hydrogen and a nonmetal; e.g.: CH₄ (methane), SiH₄ (silane).

Saline H.: A compound of hydrogen and a strongly electropositive metal; e.g.: Li⁺, Na⁺, K⁺, Cs⁺, Fr⁺, Ca⁺, Sr⁺, Ba⁺ in which hydrogen is present as H.

Hydrocarbons: (Gk. hydro, water; L. carbo, charcoal) Organic compound consisting of H and C atoms only.

Aliphatic HC: Hydrocarbons that do not contain the benzene group or the benzene ring.

Aromatic HC: A hydrocarbon that contains one or more benzene rings.

Unsaturated HC: Hydrocarbons that contain carbon-carbon double bonds or carbon-carbon triple bonds.

Ion: An atom or molecule that has lost or gained one or more electrons, and thus becomes positively or negatively charged; i.e.: Al³⁺ (mono-atomic ion), SO₄⁻ (poly-atomic ion); see also chemistry-atom.

Anion: An ion with a net negative charge, i.e.: F⁻, SO₄⁻, etc; see table below.

Cation: An ion with a net positive charge, i.e.: Na⁺, NH₄⁺, Al³⁺, etc; see table below.

Ionic Compound: Any neutral compound containing cations and anions.

Ion Pair: A species made up of at least 1 cation and at least 1 anion held together by electrostatic forces.

Polarized I.: The distorted electron cloud of an ion (or atom); see also chemistry atom; e.g.: Anion (>cloud; >charge): I⁻ (r = 220pm) easier distortable than F⁻ (r = 133pm); S²⁻ easier polarizable than Cl⁻.

Cation (<cloud; <charge): Li²⁺ more covalent than Cs⁺; Be²⁺ more covalent then cations of 4th period.

Lewis:

L. **Formula:** (for ionic compounds) A representation of the structure of an ionic compound, showing the formula unit of ions in terms of their Lewis diagram.

L. **Diagram:** (for atoms and ions) The chemical symbol of an element, with a dot for each valence electron.

L. **Dot Symbol:** The symbol of an element with one or more dots that represent the number of valence electrons in an atom of the element.

L. **Resonance:** A blending of Lewis structures into a single composite, hybrid structure; see resonance; e.g.: $\text{O}=\text{S}=\text{O} \leftrightarrow \text{O}=\text{S}-\text{O}$

L. **Structure:** A representation of covalent bonding using Lewis dot symbols. Shared electron pairs are shown either as lines or as pairs of dots between two atoms, and lone pairs are shown as pairs of dots on individual atoms.

Ligand: A molecule or an ion that is bonded to the metal ion in a complex ion.

Lone Pairs: Valence electrons that are not involved in covalent bond formation. LP-electrons occupy more space and experience a greater repulsion from neighboring lone pairs than bonding pairs; see valence - VSEPR.

Mesomerism: see resonance.

Molecule: 1) Smallest possible unit of a compound, that possesses the chemical properties of the compound. 2)

A definite and distinct, electrically neutral group of bonded atoms; i.e.: H_2 , NH_3 , CH_3COOH , etc.

M. **Equation:** Equations in which the formulas of the compounds are written as though all species existed as molecules or whole units; e.g.: $\text{C}_6\text{H}_{12}\text{O}_6$, $\text{C}_2\text{H}_3\text{CO}_2\text{H}$, etc.

M. **Orbital:** see orbital.

M. **Weight:** The relative weight of a molecule when the weight of the most frequent kind of carbon atom (its isomers) is taken as 12; the sum of the relative weights of the atoms in a molecule;

e.g.: $M_{\text{H}_2\text{O}} = 2 \cdot \text{H} + \text{O} = 18.016$ [g/mol]

Diatomic M.: A molecule that consists of two atoms; i.e.: H_2 , CO , etc.

Homonuclear Diatomic M.: Diatomic molecules containing atoms of the same element; e.g.: O_2 , H_2 , N_2 , etc.

Polar M.: A molecule that possess a dipole moment; see there.

Octet Rule: An atom other than hydrogen and helium that tends to form bonds until it is surrounded by eight valence electrons by sharing or transferring them - see also chemistry - stoichiometry.

$$B_x = \frac{1}{2} \cdot (\sum V_e + \sum V_8)$$

e.g.: N_2 (2N) has the equivalent of $10e^-$, $16e^-$ are possible to

obtain noble gas configuration: $16 - 10 = 6 \equiv 3$ covalent bonds

($:\text{N} \equiv \text{N}:$); elements from the 2nd period can't form more than 4 covalent bonds.

Exception to OR.: 1) Rule valid for covalent bonds only; 2) there are irregularities from the 3rd period of the periodic table onwards, since d-orbitals provide extra electron, seemingly interfering with the octet rule.

Orbital: The idealized representation of electron moving around its nucleus.

Hybrid O.: Atomic orbitals obtained when two or more nonequivalent orbitals of the same atom combine; a mixed orbital is obtained by blending together atomic orbitals on the same atom;

e.g.: sp^3 hybrid orbital of carbon in CH_4 : to make 4 covalent bonds for H possible the outer bonding shell in the C atom has to be altered in the following way:

$\text{C} = 2s^2 2p^2 = 2s(\uparrow\downarrow) + 2p(\uparrow, \uparrow, _) \rightarrow 2sp^3(\uparrow, \uparrow, \uparrow, \uparrow)$ with H having only a 1s orbiting electron it connects via a σ -bond to the hybridized C-atom to give rise to a tetrahedrally shaped molecule.

Molecular O.: An orbital that results from interaction of the atomic orbitals of the covalently bonding atoms; it spreads over all the atoms in a molecule. The formation of a bonding molecular orbital caused by the orientation of electron spin within the orbital; e.g.: H-H atom: $1s\sigma^*$ = antibonding (repelling, opposite spin); $1s\sigma$ = bonding (attraction, same spin);

i) every molecular orbital can accept two electrons;

i) bonding molecular orbitals are energetically far more stable than antibonding MO's;

e.g.: $E_{1s\sigma} < E_{1s\sigma^*}$; or $E_{2s\sigma} < E_{2s\sigma^*} < E_{2p_x\sigma} < E_{2p_y\pi} < E_{2p_z\pi} < E_{2p_y\pi^*} < E_{2p_x\sigma^*}$

i) orbitals are filled with electrons after the Aufbau principle, see atom;

i) the overall number of molecular orbitals consists of the sum of participating atomic orbitals.

Delocalized Molecular O.: Molecular orbitals that are not confined between two adjacent bonding atoms but actually extend over three or more atoms, or even spread over the entire molecule; e.g.: π -bonding.

- **Antibonding MO.:** A molecular orbital that is of higher energy and lower stability than the atomic orbital from which it was formed; energy has to be expended to build this molecule; probability of finding an electron within the boundary surface of the two atomic orbitals is low - cancellation in the internuclear region due to opposite spin $-\frac{1}{2}(\downarrow)$ & $+\frac{1}{2}(\uparrow)$;
- **Bonding MO.:** A molecular orbital that is of lower energy and greater stability than the atomic orbitals from which it was formed (energy will be released), hence more stable; (probability of finding an electron within the boundary surface of the two atomic orbitals is high - addition in the internuclear region occurs due to same spin $-\frac{1}{2}(\downarrow)$ & $-\frac{1}{2}(\downarrow)$ or $+\frac{1}{2}(\uparrow)$ & $+\frac{1}{2}(\uparrow)$).

Covalent Bond.: A bond in which two electrons are shared by two atoms; see bond π -, σ - bond.

Bond Order: The difference between the numbers of electrons in bonding molecular orbitals and antibonding molecular orbitals, divided by two;

e.g.: $H_2 = \frac{1}{2}(2-0) = 1$; bonding dominate over antibonding electrons, therefore H_2 is a valid molecule.

e.g.: $He_2 = \frac{1}{2}(2-2) = 0$; bonding and antibonding electrons cancel each other, He_2 is inexistant molecule.

Resonance (mesomerism): The use of two or more Lewis structures to represent a particular molecule (deducted from its formal charge - see there).

R. **Structure:** One of two or more alternative Lewis structures for a single molecule that cannot be described fully with a single Lewis structure; see also Lewis resonance.

Stereoisomerism: The occurrence of two or more compounds with the same types and numbers of atoms and the same chemical bonds but different spatial arrangements; e.g. isomers like glucose and fructose = $C_6H_{12}O_6$

Valence: The number of bonds that an atom can form.

V. **Band:** The full band of orbitals in a solid.

V. **Bond Theory:** The description of bond formation in terms of the pairing of spins in the atomic orbitals of neighboring atoms.

V. **Electron:** The outer e of an atom, involved in chemical bonding; i.e.: n = 2 shell of period 2 atoms.

V. **Shell:** The outermost electron-occupied shell of an atom, usually needed in bonding.

V. **Shell-Electron-Pair Repulsion Model (VSEPR):** A theory for predicting the shapes of molecules, using the fact that electron pairs repel one another.

- electron pairs repel another, hence orient themselves in a manner with the least obstruction;
Lone pair versus lone pair repulsion > LP versus bonding pair repulsion > BP versus BP repulsion
- the molecular shape is determined by the position of the central atoms only

CO_2 or C_2H_2 :	2 pointed \rightarrow linear;
BF_3 or SO_3 or CO_3^-	3 pointed \rightarrow trigonal-planar,
CH_4 or BF_3 or H_3PO_3 :	4 pointed \rightarrow tetrahedral;
PF_5 or SO_3 :	5 pointed \rightarrow trigonal-bipyramidal;
SF_6 or SeF_6 :	6 pointed \rightarrow octahedral;
IF_7 :	7 pointed \rightarrow pentagonal-bipyramidal;

V. **Shell Expansion:** The use of d orbitals in addition to s and p orbitals to form a covalent bond.

Intermolecular attraction (p = 1atm):

*) Hydrogen bonding

Molecule	Dipole Moment [C·m] = [D]	Attracting force [kJ/mol]		Melting point [K]	Boiling point [K]
		Dipol-dipol F. F.	London		
CO	0.4	0.0004	8.74	74	82
HI	1.3	0.025	27.9	222	238
HBr	2.6	0.69	21.9	185	206
HCl	3.4	3.31	16.8	158	188
NH_3	5.0	13.3	14.7	195*	240*
H_2O	6.1	36.4	9.0	273*	373*

Common ionic compounds:

Substance	Notation	Substance	Notation	Substance	Notation
Azide	N_3^-	Hydroxide	OH^-	Nitrate	NO_3^-
Ammonium	NH_4^+	Hydrogen Carbonate	HCO_3^-	Nitride	N^{3-}
Bromide	Br^-	Hydrogen Phosphate	HPO_4^-	Nitrite	NO_2^-
Carbonate	CO_3^{2-}	Hydronium	H_3O^+	Oxide	O^{2-}
Chlorate	ClO_3^-	Di H ydrogen Phosphate	$H_2PO_4^-$	Oxonium	OH_3^+
Per Ch lorate*	ClO_4^-	Hydrogen Sulfate	HSO_3^-	Per O xide*	O_2^{2-}
Chloride	Cl^-	Iodide	I^-	Phosphate	PO_4^{3-}
Chlorite	ClO_2^-	Iron-II ion	Fe^{2+}	Phosphide	P^{3-}
Hypo Ch lorite	ClO^-	Iron-III ion	Fe^{3+}	Phosphonium	PH_4^+
Cobalt-II ion	Co^{2+}	Lead-II ion	Pb^{2+}	Sulfate	SO_4^{2-}
Cobalt-III ion	Co^{3+}	Lead-IV ion	Pb^{4+}	Sulfide	S^{2-}
Copper-I ion	Cu^+	Manganese-II ion	Mn^{2+}	Sulfite	SO_3^{2-}
Copper-II ion	Cu^{2+}	Manganese-III ion	Mn^{3+}	Tin-II ion	Sn^{2+}
Cyanide	CN^-	Mercury-I ion	Hg_2^{2+}	Tin-IV ion	Sn^{4+}
Fluoride	F^-	Mercury-II ion	Hg^{2+}		

(*) per (L. allover) the elements max. ability to combine with O; i.e.: highest O-content

Relative electronegativity of elements from the main groups [V]

H 2.2								He -
Li 1.0	Be 1.6		B 2.0	C 2.6	N 3.0	O 3.4	F 4.0	Ne -
Na 0.9	Mg 1.3		Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2	Ar -
K 0.8	Ca 1.0		Ga 1.8	Ge 2.0	As 2.2	Se 2.6	Br 3.0	Kr -
Rb 0.8	Sr 0.9		In 1.8	Sn 2.0	Sb 2.1	Te 2.1	I 2.7	Xe -
Cs 0.8	Ba 0.9		Ti 2.0	Pb 2.3	Bi 2.0	Po 2.0	At 2.2	Rn -

Common molecular compounds

Compound	Notation	Compound	Notation	Compound	Notation
Acetaldehyde	CH ₃ CHO	Cyclo Hexane	C ₆ H ₁₂	Octane	C ₈ H ₁₈
Acetic Acid	CH ₃ COOH	Hexane		Oxalic Acid	(COOH) ₂
Acetylene	C ₂ H ₂			Oxygen	O ₂
Acetone	CH ₃ COCH ₃	Hydrazine	N ₂ H ₄	Ozone	O ₃
Aniline	C ₆ H ₅ NH ₂	Hydrobromic Acid	HBr	Phenol	C ₆ H ₅ OH
Aluminium Dioxide	Al ₂ O ₃	Hydrobromous Acid	HBrO	Phosphine	PH ₃
Ammonia	NH ₃			Phosphoric Acid	H ₃ PO ₄
Ammonium Chlorate	NH ₄ ClO ₃	Hydrochloric Acid	HCL	Phosphorous Acid	H ₃ PO ₃
Ascorbic Acid	C ₆ H ₈ O ₆	Hydrochlorous Acid	HClO	Phosphorous Pentachloride	PCl ₅
Barium Chloride Dihydrate	BaCl·2H ₂ O			Phosphorous Trichloride	PCl ₃
Barium Hydroxide	Ba(OH) ₃	Hydrobromic Acid	HBr	Tetra Phosphorous Decaoxide	P ₄ O ₁₀
Barium Peroxide	BaO ₂	Hydrocyanic Acid	HCN	Cyclo Propane	C ₃ H ₆
Benzene	C ₆ H ₆	Hydroflouric Acid	HF	Propane	C ₃ H ₈
Benzoic Acid	CH ₃ CO ₂ H	Hydrogen Bromide	HBr	Propylene	C ₃ H ₆
Butane	C ₄ H ₁₀	Hydrogen Peroxide	H ₂ O ₂	Potassium Dichromate	K ₂ Cr ₂ O ₇
DiBorane	B ₂ H ₆			Potass. Dihydrogenphosphate	KH ₂ PO ₄
Borone Triflouride	BF ₃	Hydrogen Sulfide	H ₂ S	Potassium Hydroxide	KOH
Per Bromic Acid	HBrO ₄	Hydroiodic Acid	HI	Potassium Permanganate	KMnO ₄
Calcium Carbonate	CaCO ₃	Hydrosulfuric Acid	H ₂ S		
Calcium Oxide	CaO			Silane	SiH ₄
Calcium Phosphate	Ca ₃ (PO ₄) ₂	Iodic Acid	HIO ₃	Silicon Carbide	SiC
		Iodine Heptaflouride	IF ₇	Silicon Tetrachloride	SiCl ₄
Carbon Dioxide	CO ₂	Di Iodine Pentaoxide	I ₂ O ₅	Di Silicon Hexabromide	Si ₂ Br ₂
Carbon Disulfide	CS ₂	Per Iodic Acid	HIO ₄	Sodiumchloride	NaCl
Carbon Monoxide	CO			Sodiumfloride	NaF
		Isopropanol	C ₃ H ₈ O	Sodium Hydroxide	NaOH
Cesium Sulfide	Cs ₂ S	Lead-II Oxide	PbO	Sodium Peroxide	Na ₂ O ₂
Chloric Acid	HClO ₃	Lithiumcarbonate	Li ₂ CO ₃	Strontium Nitrate Tetrahydrate	Sr(NO ₃)·7H ₂ O
Chlorine Triflouride	ClF ₃	Lithium Chloride Monohydrate	LiCl·H ₂ O		
Di Chloride Heptaoxide	Cl ₂ O ₇	Lithiumsulfide	Li ₂ SO ₃	Sucrose	C ₁₂ H ₂₂ O ₁₂
Per Chloric Acid	HClO ₄				
Chloroform	CHCl ₃	Magnesium Sulfate Heptahydrate	MgSO ₄ ·7H ₂ O	Sulfuric Acid	H ₂ SO ₄
Chlorous Acid	HClO ₂			Sulfur Dichloride	SCl ₂
Copper-II Bromide	CuBr ₂	Mercury-I Nitrite	Hg ₂ (NO ₂) ₂	Sulfur Dioxide	SO ₂
Copper-II Nitrite	Cu(NO ₂) ₂			Sulfur Hexaflouride	SF ₆
Copper Sulfate Pentahydrate	CuSO ₄ ·5H ₂ O	Methane	CH ₄	Sulfur Trioxide	SO ₃
Ethanol	C ₂ H ₅ OH	Methanol	CH ₃ OH	Tetra Sulfur Tetranitride	S ₄ N ₄
Ethane	C ₂ H ₆	Methylamine	CH ₃ NH ₂	Sulfurous Acid	H ₂ SO ₃
Ethylene	C ₂ H ₄				
Formaldehyde	HCHO	Nitric Acid	HNO ₃	Toluene	C ₇ H ₈
Formic Acid	HCO ₂ H	Nitrogen Dioxide	NO ₂	Urea	CO(NH ₂) ₂
Fructose	C ₆ H ₁₂ O ₆	Nitrogen Trichloride	NCl ₃	Water	H ₂ O
Glucose	C ₆ H ₁₂ O ₆	Di Nitrogen Pentaoxide	N ₂ O ₅		
Glycine	NH ₂ CH ₂ CO ₂ H	Di Nitrogen Tetraoxide	N ₂ O ₄		

Glossary - Chemistry (Stoichiometry)

Adsorption: The process of binding a substance to a surface by releasing energy; see also catalyst.

Chemical A.: Molecules are bound chemically via lone pairs (σ , bonds, covalent-like forces) to the surface; e.g.: catalytic hydration of olefins.

Physical A.: Molecules are bound via dispersion forces to the surface (releases less energy than in chemically adsorbed processes; e.g.: charcoal filter).

Alcohol: An organic compound containing the hydroxyl group -OH.

Aldehydes: Compounds with a carbonyl functional group and the general formula RCHO, where R is an H-atom, an alkyl, or an aryl group.

Allotrops: Two or more forms of the same element that differ significantly in chemical and physical properties.

Alloy: A solid solution composed of two or more metals, or of a metal or metals with one or two nonmetals.

Amalgam: An alloy of mercury with another metal or metals.

Amines: Organic bases that have the functional group -NR_2 , where R may be H, an alkyl group, or an aryl group.

Arrhenius Behavior: see kinetics.

Aryl Group: A group of atoms equivalent to a benzene ring or a set of fused benzene rings, less one H-atom.

N_A - **Avogadro's Number:** The number of particles in a mole: $6.02252 \cdot 10^{23}$ [particles/mole]

Catalyst: A substance that increases the rate of a reaction without being consumed in the reaction; activation energy E_A is lowered significantly (increases rate of reaction) once catalyst is used in process without changing the overall energy of reaction (ΔH);

e.g.: $\text{N}_2\text{O}(\text{g}) \rightarrow (\text{Au}) \rightarrow \text{N}_2\text{O}(\text{docked onto Au-cat})$

$\text{N}_2\text{O}(\text{docked onto Au-cat}) \rightarrow \text{N}_2(\text{g}) + \text{O}(\text{docked onto Au-cat})$

$2\text{O}(\text{docked onto Au-cat}) \rightarrow \text{O}_2(\text{g})$ and so forth

Homogeneous C.: A catalyst is homogeneous if it is present in the same phase as the reactants;

e.g.: $\text{Br}_2(\text{aq})$ for the decomposition of $\text{H}_2\text{O}_2(\text{aq})$;

Heterogeneous C.: A catalyst is heterogeneous if it is in a different phase from the reactant;

e.g.: Pt/Rh(s) for the decomposition of NO and CO (Ostwald process)

Chain Reaction: Is a reaction in which an intermediate reacts to produce another intermediate - see radicals, or physics nuclear decay.

Chemical Equation: see equation.

Compound: A substance composed of atoms of two or more elements chemically united in fixed properties.

Concentration Units: The following equations are commonly used when operating with liquid mixtures.

β - **Mass Concentration:** The mass of solute per liter of solution:

$$\beta_{(x)} = m_{(x)} / V_{(\text{Soln})} \quad [\text{g/l}]$$

$m_{(x)}$, mass [g]
 $V_{(\text{Soln})}$, volume of solution [l]

w - **Mass Percent Composition** (weight percentage): The percent by mass of one solvent over the total mass of solvent and solution:

e.g.: 30% NaCl in 70% water

$$w_{(A)} = 100 \cdot m_{(A)} / (m_{(A)} + m_{(B)} + \text{etc.}) \quad [\%]$$

b - **Molality:** The number of moles of solute dissolved in one kilogram of solvent; b is not affected by temperature change:

$$b_{(x)} = n_{(x)} / m_{(\text{Svt})} \quad [\text{mol/g}]$$

$m_{(A)}$, mass of solute A [kg]
 $n_{(x)}$, molar amount [mol]
 $m_{(\text{Svt})}$, mass of solvent [g]

Remember: A solution with a given molality always maintains the same molar with whatever solute, as long as the solvent has not changed!

n - **Molar Amount:** The amount of an element per molar mass:

$$n_{(x)} = m_{(x)} / M_{(x)} \quad [\text{mol}]$$

$m_{(x)}$, mass [g]
 $M_{(x)}$, molar mass [g/mol]
 $n_{(x)}$, molar amount [mol]
 $V_{(\text{Soln})}$, volume of solution [l]

c - **Molar Concentration** (M - molarity): The number of moles of solute in one liter of solution (c increases with temperature):

$$c_{(x)} = n_{(x)} / V_{(\text{Soln})} \quad [\text{mol/l}]$$

x - **Molar Fraction:** The ratio by mole of one solvent over the sum of moles of all components:

$$x_{(A)} = n_{(A)} / (n_{(A)} + n_{(B)} + n_{(C)} + \text{etc.}) \quad [-]$$

Molar Solubility: The number of moles of solute in one liter of a saturated solution (mol/L) = c.

Equation: An expression showing the chemical formulas of the reactants and products (both in symbols).

Ionic EQ.: An equation that shows dissolved ionic compounds in terms of their free ions;

i.e.: $\text{Ag}^+(\text{aq}) + \text{NO}_3^-(\text{aq}) + \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{Na}^+(\text{aq}) + \text{NO}_3^-(\text{aq})$

Net Ionic EQ.: The equation showing the net change, obtained by canceling the spectator ions in an ionic equation; i.e.: $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$

Writing ionic and net ionic EQ: 1) Write a balanced molecular EQ for the reaction; 2) Rewrite the equation to indicate which substances are in ionic form in solution (all electrolyte in solution dissociate into anions and cations; group-I elements); 3) Identify and cancel spectator ions (appear on both sides of the EQ) to arrive at the net ionic EQ;

Balanced EQ.: A chemical equation in which the same number of atoms of each element appear on both sides of the equation; i.e.: $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$

Skeletal EQ.: An unbalanced equation that summarizes the qualitative information about the reaction; i.e.: $\text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O}$,

Product: A substance formed in a chemical equation.

Reactant: A starting material in a chemical reaction; a reagent taking part in a specified reaction.

Symbol: One- or two-letter abbreviation of an element's name.

Equilibrium: A state in which there are no observable changes as time goes by.

K_C - E. Constant: A number equal to the ratio of the equilibrium concentration of gaseous products to the equilibrium concentrations of gaseous reactants, each raised to the power of its stoichiometric coefficient (ignored when in their solid or liquid state); by convention, numerator stands for products, and denominator stands for reactants; **K_C is temperature dependent;**

any solid that precipitates or compound that liquefies, is left out.

k_F, p_F, k_R, p_R , rate c. of forward / reverse reactions [mol/(l·s)]

e.g.: $aA + bB \leftrightarrow cC + dD$

$K_C = k_F/k_R = c_{(C)}^c \cdot c_{(D)}^d / (c_{(A)}^a \cdot c_{(B)}^b)$ [var]

$K_C > 10^3$: favors products strongly

$K_P = p_F/p_R = p_{(C)}^c \cdot p_{(D)}^d / (p_{(A)}^a \cdot p_{(B)}^b)$ [var]

$K_C > 10^{-3} - 10^3$: reactants & products at equilibrium.

where $K_P = K_C \cdot (R \cdot T)^{\Delta n}$

$K_C < 10^{-3}$: favors reactants strongly

Δn , moles of gaseous products - moles of gaseous reactants

K_C , EC for concentrations

Q_C - Reaction Quotient: The ratio of the product of the concentrations of the products to that of the reactants (defined like the equilibrium constant) at an arbitrary stage of reaction;

K_P , EC for partial pressures

e.g.: $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$

c , molar concentration [mol/l]

$Q_C < K_C$: reverse-reaction towards products (shift to right)

p , partial pressure [N/m²]

$Q_C = K_C$: system in equilibrium

$\Delta n = (c+d) - (a+b)$ [mol]

$Q_C > K_C$: forward-reaction towards reactants (shift to left)

$K_C = c_{(\text{NH}_3)}^2 / (c_{(\text{N}_2)} \cdot c_{(\text{H}_2)}^3)$

$Q_C = c_{0(\text{NH}_3)}^2 / (c_{0(\text{N}_2)} \cdot c_{0(\text{H}_2)}^3)$

$c_{0(X)}$, initial arbitrary c. [mol/l]

E. Vapor Pressure: The vapor pressure measured under

dynamic equilibrium of condensation and evaporation; see chemistry liquid and physics - heat.

Dynamic E.: The condition in which a forward process and its reverse are occurring simultaneously at equal rates; e.g.: vaporizing and condensing; chemical reactions at equilibrium, etc.; e.g.:

$\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \leftrightarrow 2\text{HI}(\text{g})$;

v_F, v_R , rate of forward/reverse

$v_F = k_F \cdot c_{(\text{H}_2)} \cdot c_{(\text{I}_2)}$; $v_R = k_R \cdot c_{(\text{HI})}^2$;

k_F, k_R , rate constants [1/s]

$v_F = v_R$ in equilibrium: $c_{(\text{HI})}^2 / (c_{(\text{H}_2)} \cdot c_{(\text{I}_2)}) = k_F/k_R = K$

$c_{(X)}$, molar concentration [mol/l]

Heterogeneous E.: An equilibrium state in which the reacting species are not all in the same phase.

e.g.: $\text{CaCO}_3(\text{s}) \leftrightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$

$K_C = c_{(\text{CaO})} \cdot c_{(\text{CO}_2)} / c_{(\text{CaCO}_3)}$

solid substance has to be converted from density [g/cm³] to molar concentration [mol/l].

Homogeneous E.: An equilibrium state in which all reacting species are in the same phase.

e.g.: $\text{N}_2\text{O}_4(\text{g}) \leftrightarrow 2\text{NO}_2(\text{g})$

$K_P = p_{(\text{NO}_2)}^2 / p_{(\text{N}_2\text{O}_4)}$

Multiple E.: If a reaction can be expressed as the sum of two or more reactions, the equilibrium constant for the overall reaction is given by the product of the equilibrium constants of the individual reactions; e.g.:

$\text{H}_2\text{CO}_3(\text{aq}) \leftrightarrow \text{H}^+(\text{aq}) + \text{HCO}_3^-(\text{aq})$

$K_1 = c_{(\text{H}^+)} \cdot c_{(\text{HCO}_3^-)} / c_{(\text{H}_2\text{CO}_3)}$

$\text{HCO}_3^-(\text{aq}) \leftrightarrow \text{H}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq})$

$K_2 = c_{(\text{H}^+)} \cdot c_{(\text{CO}_3^{2-})} / c_{(\text{HCO}_3^-)}$

$\text{H}_2\text{CO}_3(\text{aq}) \leftrightarrow 2\text{H}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq})$

$K_C = K_1 \cdot K_2$

$K_C = c_{(\text{H}^+)}^2 \cdot c_{(\text{CO}_3^{2-})} / c_{(\text{H}_2\text{CO}_3)}$

Law of Mass Action: For an equilibrium of the form $aA + bB \leftrightarrow cC + dD$, the reaction quotient

$Q_C = c_{(C)}^c \cdot c_{(D)}^d / c_{(A)}^a \cdot c_{(B)}^b$; evaluated by using the equilibrium molar concentrations of the reactants and products, is equal to a constant K_C which has a specific value for a given reaction and temperature.

Van't Hoff Equation: It shows how the equilibrium constant K_C varies with temperature.

T , temperature [K]

$\ln(K_{C2}/K_{C1}) = -(1/T_2 - 1/T_1) \cdot \Delta H^\circ/R$

K_C , equilibrium constant [var]

Functional Group: That part of a molecule characterized by a special arrangement of atoms that is largely responsible for the chemical [kJ/mol]

R , gas constant 8.314 [kJ/mol]

behavior of the parent molecule.

ΔH° ; std enthalpy of react.

Formula (Chemical F.): A collection of chemical symbols and subscripts showing the composition of a substance.

Empirical F.: A chemical formula that shows the relative numbers of atoms of each element in a compound, i.e.: CH for acetylene (C₂H₂) or benzene (C₆H₆), etc.

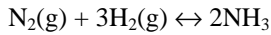
Molecular F.: A combination of chemical symbols and subscripts showing the actual numbers of atoms of each element present in a molecule; i.e.: H₂O, SF₆, C₆H₁₂O₆; etc.

Structural F.: A chemical formula that shows the groupings of atoms in a compound.

Lewis Dot Symbol: The symbol of an element with one or more dots that represent the number of valence electrons in an atom of the element.

Lewis Structure: A diagram showing how electron pairs are shared between atoms in a molecule; e.g.: H-Cl::: or ::O=C=O::

Haber Bosch Process: The catalyzed synthesis (Pt) of ammonia at high pressure and high temperature:

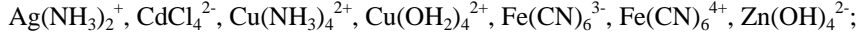


Half Life: see kinetics.

Hydrogenation: The addition of hydrogen, especially to compounds with double and triple-carbon bonds.

Ion: (Gk, to go) An atom or molecule that has lost or gained one or more electrons, and thus becomes positively or negatively charged; i.e.: Al³⁺ (mono-atomic ion), SO₄⁻ (poly-atomic ion); see chemistry atom & molecule.

Complex I.: An anion containing a central metal cation bonded to one or more molecules or ions; e.g.:



Spectator I.: Ions that are present in the reaction but remain unchanged (compare net ionic w/ ionic eq.)

Kinetics: The speed, or rate, at which a chemical reaction occurs.

E_a, activation energy [J/mol]

Arrhenius Behavior: A reaction shows Arrhenius behavior

k, rate constants [1/s]

if a plot of ln(k) against 1/T is a straight line.

A, collision frequency [-]

e.g.: $k = A \cdot e^{-E_a/(R \cdot T)}$; $\ln(k) = \ln(A) - (E_a/(R \cdot T))$

e, eulers number 2.7183 [-]

for a given reaction at 2 different T's:

R, gas constant 8.314 [J/mol]

$$\ln(k_2/k_1) = E_a \cdot (T_2 - T_1) / (R \cdot T_1 \cdot T_2)$$

T, temperature [K]

K. Theory (kinetic molecular theory); A theory of properties of an ideal gas in which point-like molecules are in continuous random motion and there are no interactions between them; see chemistry gas.

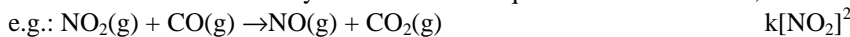
t_{1/2} - Half Life: The time required for the concentration of a reactant to decrease to half of its initial concentration; see O, 1st, 2nd order reactions.

Kinetic Reactions:

k - KR. Constant: The constant of proportionality in a rate law; k is dependent upon the temperature and can only be determined experimentally; e.g.: c_(A) at t₁ - c_(A) at t₂ / (t₂ - t₁)

k = decrease of c(reactants)/t [1/(mol·s)] or [1/s] or [mol/(l·s)] depending upon the rate order.

Rate Law: An EQ expressing the instantaneous reaction rate in terms of the concentrations, at that instant, of the substances taking part in the reaction; it can only be determined experimentally; it cannot be written down from the stoichiometry of the chemical equation for the reaction;



KR. Order: The power to which the concentration of a single substance is raised in a rate law;

e.g.: k[SO₂]¹·[SO₃]^{-1/2}, then the reaction is 1st order in SO₂ and of order -1/2 in SO₃ and +1/2 overall.

- **Pseudo 1st Order KR.:** A reaction with a rate law that is effectively first order because one substance has a virtually constant concentration.

- **0 Order KR.:** A reaction with a rate that is independent of the concentration of the reactant; i.e.: the catalyzed decomposition of ammonia; plot of c_A is linear over t:

$$\text{e.g.: } v_{(\text{AX})} = -k \quad [\text{mol}/(\text{l} \cdot \text{s})] \quad c_{(\text{A})} = -k \cdot t + c_{(\text{A}0)} \quad [\text{mol}/\text{l}] \quad t_{1/2} = c_{(\text{A}0)} / (2 \cdot k) \quad [\text{s}]$$

- **1st Order KR.:** A reaction whose rate depends on reactant concentration raised to the first power; (directly proportional to the concentration); plot of ln(c_A) is linear over t:

$$\text{e.g.: } v_{(\text{AX})} = -k \cdot c_{(\text{AX})}; \quad [1/\text{s}] \cdot [\text{mol}/\text{l}] \quad c_{(\text{A})} = c_{(\text{A}0)} \cdot e^{-k \cdot t} \quad [\text{mol}/\text{l}] \quad t_{1/2} = \ln(2) / k \quad [\text{s}]$$

- **2nd Order KR.:** A reaction whose rate depends on reactant concentration raised to the second power or on the concentration of two different reactants, each raised to the first power (directly proportional to the square of the concentration); plot of 1/c_A is linear over t:

$$\text{e.g.: } v_{(\text{AX})} = k \cdot c_{(\text{AX})}^2 \quad [1/(\text{mol} \cdot \text{s})] \cdot [\text{mol}^2/\text{l}^2] \quad 1/c_{(\text{A})} = k \cdot t + 1/c_{(\text{A}0)} \quad [\text{mol}/\text{l}] \quad t_{1/2} = 1 / (k \cdot c_{(\text{A}0)}) \quad [\text{s}]$$

v_(AX) - KR. Rate: The change in concentration of a substance divided by the time it takes for the change to occur: an increase in product automatically implies a decrease in reactant and vice versa;

e.g.: v_(AX) = dc_(AX)/dt [mol/(l·s)] with v_(AX) positive = increase; v_(AX) negative = decrease.

KR. Sequence: A series of reactions in which products of one reaction take part as reactants in the next;



- **Mono-Molecular KRS.:** A single molecule forms the product; v = k·c_(A); e.g.: radioactive decay

- **Di-Molecular KRS.:** Involves two molecules; v = k·c_(A)·c_(X); with A and X 1st order, with the overall reaction itself as 2nd order.

Law of Definite Proportions: Different samples of the same compound always contain its constituent elements in the same proportions by mass.

Law of Mass Action: see equilibrium.

Law of Multiple Proportions: If two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element are in ratios of small whole numbers.

Le Chatelier's Principle: If an external stress is applied to a system at equilibrium, the system will adjust itself in such a way as to partially offset the stress; i.e.: a reaction at equilibrium tends to proceed in the endothermic reaction when the temperature is raised.

- Change of concentration: When the concentration of any of the reactants or products at equilibrium are changed, the position of the equilibrium shifts so as to reduce the change in concentration that was made, with the net effect that K_C remains unchanged.
- Change of pressure and volume: i) In gaseous substances with $\Delta n \neq 0$ changes of pressure and volume do influence the equilibrium; e.g.: The amount of product (NH_3) at equilibrium increases with rising pressure; $\text{N}_2 + 3\text{H}_2 \leftrightarrow 2\text{NH}_3 + \text{heat}$ since the formation of two molecules of NH_3 causes the disappearance of four molecules of reactants ($\text{N}_2 + 3\text{H}_2$), thus decreasing the total number of molecules.
 - i) In reactions involving liquids, moderate pressure changes do not affect the equilibrium;
 - i) In reactions involving gaseous substance with $\Delta n = 0$, changes of pressure will not affect the equilibrium;
- Change of Temperature: In gaseous substances a temperature increase favors an endothermic reaction, and a temperature decrease favors exothermic reactions:
 - T-increase ($T_2 < T_1$): exothermic reaction: $\Delta H^\circ < 0$; $K_2/K_1 < 1$ or $K_2 < K_1$
 endothermic reaction: $\Delta H^\circ > 0$; $K_2/K_1 > 1$ or $K_2 > K_1$; equilibrium concentration of products increases; position of equilibrium shifts to right; K increases;
 - T-decrease ($T_2 < T_1$): endothermic reaction: $\Delta H^\circ > 0$; $K_2/K_1 < 1$ or $K_2 < K_1$;
 exothermic reaction: $\Delta H^\circ < 0$; $K_2/K_1 > 1$ or $K_2 > K_1$; equilibrium concentration of products decreases; position of equilibrium shifts to left; K decreases;

Lewis: Lewis dot symbol and Lewis structure see formula.

Limiting Reagent: The reactant used up in a reaction - see yield.

Lysis: (Gk. lysis, loosening) A process of disintegration or destruction.

Glycolys.: Splitting of glucose molecules; ($\text{C}_6\text{H}_{12}\text{O}_6$) into two molecules of pyruvate, resulting in the release of energy (exothermic) in the form of two ATP molecules; does not require O_2 .

Hydrolys.: The splitting apart of two covalently bound subunits by the insertion of a molecule of water; the OH-group being incorporated into one fragment and the H-atom in the other; reverse process of synthesis occurs in digestion.

Protolys.: Dissociation of protons in an acid $\text{HA} \leftrightarrow \text{H}^+ + \text{A}^-$.

Mass: The quantity of matter in an object, or concentrated energy (see SI-units, weight).

M. Defect: The difference between the mass of an atom and the sum of the masses of its protons, neutrons, and electrons; since part of the mass is needed to hold the atom together (potential energy).

A - Mass Number: The total number of neutrons and protons present in the nucleus of an atom; i.e.: $^{14}_6\text{C}$ has 14 nucleons (6 protons and 8 neutrons).

Atomic Mass Unit (amu): A mass exactly equal to one 12^{th} of the mass of one carbon-12 atom (see molecular mass):
 $1 \text{ amu} = m_{12\text{C}} / 12 = 1.6605 \cdot 10^{-24} \text{ [g]}$

Average Atomic M.: see chemistry atom.

Law of Conservation of M.: Matter can be neither destroyed nor created.

Mass Units: The following equations are commonly used in dealing with masses in chemical equations:

β - Mass Concentration: The mass of solute per liter of solution: $m_{(x)}$, mass [g]
 $\beta_{(x)} = m_{(x)} / V_{(\text{Soln})}$ [g/l] $V_{(\text{Soln})}$, volume of solution [l]

ρ - Density: The mass of a substance divided by its volume $m_{(\text{Soln})}$, mass [g]
 $\rho_{(\text{Soln})} = m_{(\text{Soln})} / V_{(\text{Soln})}$ [g/l] $V_{(\text{Soln})}$, volume of solution [l]

M_r - Molecular M.: The sum of the atomic masses (in amu) present in the molecule;
 i.e.: $M_r(\text{H}_2\text{O}) = 2 \cdot 1.008$ (atomic mass of H) + 16.00 (atomic mass) of O = 18.02 [amu].

n - Molar Amount: The amount of an element per molar mass: $m_{(x)}$, mass [g]
 $n_{(x)} = m_{(x)} / M_{(x)}$ [mol] $M_{(x)}$, molar mass [g/mol]

M - Molar Mass: The relative mass (g, or kg) per mole of atoms (amu), molecules, or other particles - see also atom-mass: m_{av} , average mass [amu]

$M_{(x)} = m_{\text{av}} \cdot N_A$ [g/mol] N_A , Avogadro's c. $6.022 \cdot 10^{23}$ [atoms/mol]

i.e.: $M(\text{H}_2\text{O}) = 2 \cdot 1.008$ [g/mol] of H + 16.00 [g/mol] of O = 18.02 [g/mol];

Molar Mass Fraction: Ratio of the number of moles of one component of a mixture to the total number of moles of all components in the mixture; i.e.: the respective mass of H and O in a given sample of water is obtained by:

$m_{\text{H}} = m_{\text{H}_2\text{O}} \cdot 2 \cdot M_{\text{H}} / M_{\text{H}_2\text{O}}$ [g]; $m_{\text{O}} = m_{\text{H}_2\text{O}} \cdot M_{\text{O}} / M_{\text{H}_2\text{O}}$ [g];

Molar: The quantity per mole; i.e.: molar mass (the mass per mole), molar volume (the volume per mole), etc.

b - **Molality:** see chemistry-liquids.

c - **Molar Concentration:** see chemistry-liquid.

Molar Solubility: see chemistry-liquid.

c - **Molarity:** see chemistry-liquid.

Mole: (L, massive, heap) The SI base unit for the amount of substances that contains as many elementary entities (atoms, molecules, or other particles) as there are atoms in exactly 12 grams of the carbon-12 isotope; always equal to Avogadro's number = $6.02205 \cdot 10^{23}$.

M. **Method:** The approach of determining the amount of product formed in a reaction, based on the fact that the stoichiometric coefficients in a chemical EQ can be interpreted as the number of moles of each substance; the MM consists of the following steps:

1. Write correct formula for all reactants and products and balance the resulting EQ;
2. Convert the quantities of some or all given or known substances (usually reactants) into moles;
3. Use the coefficients in the balanced EQ to calculate the number of moles of the sought or unknown quantities (usually products);
4. Using the calculated numbers of moles and the molar masses, convert the unknown quantities to whatever units are required (usually grams);
5. Check that the answer is reasonable in physical terms

Octet Rule: An atom other than hydrogen and helium that tends to form bonds until it is surrounded by eight valence electrons by sharing or transferring them - see chemistry solid.

Percent Composition: The percent by mass of each element in a compound.

Products & Reactant: see equation.

Radical: An atom, molecule, or ion with at least one unpaired electron; e.g.: NO, $\bullet\text{:O}\bullet$, $\bullet\text{CH}_3$;

R. **Reaction Mechanism:** The pathway that is proposed for an overall reactions and accounts for the experimental rate law; e.g.: $\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}$ splits into 4 sub-reactions:

- Initiation: $\text{Cl}_2 \xrightarrow{\text{blue light}} 2\text{Cl}\bullet$ blue contains more energy than red light
- Chain R.: $\text{Cl}\bullet + \text{H}_2 \rightarrow 2\text{HCl} + \text{H}\bullet$; $\text{H}\bullet + \text{Cl}_2 \rightarrow \text{HCl} + \text{Cl}\bullet$...in a closed loop
- Back R.: $\text{H}\bullet + \text{HCl} \rightarrow \text{H}_2 + \text{Cl}\bullet$
- Termination: $\text{Cl}\bullet + \text{Cl}\bullet \rightarrow \text{Cl}_2 + \text{E}$; $\text{H}\bullet + \text{H}\bullet \rightarrow \text{H}_2$ $\text{H}\bullet + \text{Cl}\bullet \rightarrow \text{HCl}$
once the light is switched off....

Reaction: A chemical change in which one substance responds to the presence of another, to a change of temperature, or to some other influence; i.e.: the sequence of elementary steps that leads to product formation.

Q_C - R. **Quotient:** see equilibrium.

Elementary R.: An individual reaction step in a reaction mechanism; e.g.: $\text{H}\bullet + \text{Cl}\bullet \rightarrow \text{HCl}$.

Reversible R.: A reaction that can occur in both directions until dynamic equilibrium is reached, see there;

e.g.: $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \leftrightarrow 2\text{HI}(\text{g})$

Classification of R.: first 5 based on the composition of products and reactants.

1. **Synthesis:** Formation of compounds from simpler starting materials; e.g.: $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$

2. **Decomposition:** Formation of simpler substances from more complex starting materials;

e.g.: $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$

3. **Replacement R.:** Exchange of partners; e.g.: $2\text{NaCl}(\text{aq}) + \text{Pb}(\text{NO}_3)_2(\text{aq}) \rightarrow 2\text{NaNO}_3(\text{aq}) + \text{PbCl}_2(\text{s})$

4. **Combustion:** Reaction with oxygen to form CO₂, H₂O, N₂, and oxides of any other elements present;

e.g.: $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$

5. **Corrosion:** Reaction of a metal with oxygen to form the metal oxide; e.g.: $4\text{Fe}(\text{s}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{Fe}_2\text{O}_3(\text{s})$

Gasevolution: Formation of gas (driving force: escape of gas);

e.g.: $\text{CaCO}_3(\text{s}) + 2\text{HCl} \rightarrow \text{CaCl}_2(\text{s}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$

Precipitation: Formation of precipitate when one solution is added to another (driving force: formation of insoluble precipitate); e.g.: $3\text{CaCl}_2(\text{aq}) + 2\text{Na}_3\text{PO}_4(\text{aq}) \rightarrow \text{Ca}_3(\text{PO}_4)_2(\text{s}) + 6\text{NaCl}(\text{aq})$

Neutralization: Reaction between any acid and a base (driving force: formation of solvent)

e.g.: $\text{HCl}(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$

RedOx: Transfer of electrons from one species to another (accompanied by atoms in many cases - driving force: e-transfer to achieve greater stability); see chemistry redox-reaction;

e.g.: $2\text{Mg}^0(\text{s}) + \text{O}_2^0(\text{g}) \rightarrow 2\text{Mg}^{+1}\text{O}^{-1}(\text{s})$

Stoichiometry: The mass relationships among reactants and products in chemical reactions.

S. **Amount:** The exact molar amount of reactants and products that appear in a balanced chemical EQ.

S. **Coefficient:** The number of moles of each substance in a chemical equation;

i.e.: 1 and 2 in: $\text{H}_2 + \text{Br}_2 \rightarrow 2\text{HBr}$

S. **Point:** The stage in a titration when exactly the right volume of solution needed to complete the reaction has been added (see chemistry - acid and base).

S. **Proportions:** Reactants in the same proportions as their coefficients in the chemical equation; i.e.: equal amounts of H_2 and Br_2 in the reaction mentioned above.

S. **Relation:** An expression that equates the relative amounts of reactants and products that participate in a reaction; i.e.: $1 \text{ mol H}_2 = 2 \text{ mol HBr}$.

Reaction S.: The quantitative relation between the amounts of reactants consumed and products formed in chemical reactions as expressed by the balanced chemical equation for the reaction.

Titration: The analysis of composition by measuring the volume of one solution (the titrant) needed to react with a given volume of another solution (the analyte).

Analyte: The solution of unknown concentration in a titration.

Titrant: The solution of known concentration added from a buret in a titration.

Van't Hoff Equation: see equilibrium.

Yield: The outcome of a chemical reaction, expressed in grams, mole, liters, etc.

Y. of **Reaction** (actual yield): The quantity of product obtained from the reaction.

Percentage Y.: The percentage of the theoretical yield of a product achieved in practice

$Y_{\%} = 100 \cdot Y_A / Y_T$ [-]: Y_A , achieved yield [g]

Theoretical Y.: The amount of product predicted by the balanced equation when all of the limiting reagent has reacted. Y_T , theoretical yield [g]

Limiting Reagent: The reactant that governs the theoretical yield of product in a given reaction.

Glossary - Chemistry (Gas)

Avogadro's Law: Equal volumes of different gases at the same temperature and pressure contain equal numbers of molecules. 1 [mol] ($6.022 \cdot 10^{23}$ molecules) of gas at $0[^\circ\text{C}]$ and 1[bar] occupies **22.414** [l] = mole-volume.

Boyle's Law: At a given temperature, the product of pressure and volume of a given mass of gas is constant; the volume of a fixed amount of gas maintained at constant temperature is inversely proportional to the gas pressure: $p_1 \cdot V_1 = \text{constant} = p_2 \cdot V_2$ [$\text{Pa} \cdot \text{m}^3$]

Colligative Properties: see chemistry - liquid.

Collision Theory: The theory of elementary gas-phase bimolecular reactions in which it is assumed that molecules react only if they collide with at least enough kinetic energy for bonds to be broken.

Activated Complex T.: A theory of reaction rates in which it is postulated that the reactants form an activated complex; i.e.: compounds temporarily formed by the reactant molecules as a result of collision.

Activated Complex: A combination of the two reactant molecules that can either go on to form products or fall apart into the unchanged reactants;

e.g.: $\text{A}_2 + \text{X}_2 \leftrightarrow \text{A}_2\text{X}_2 \rightarrow 2\text{AX}$; where A_2X_2 is very short-living.

E_A - **Activation Energy:** The minimum energy needed for reaction; the height of the activation barrier; is usually endothermic, until activated complex-level is reached, followed by a more or less exothermic reaction; resulting in a net exo-/endothermic reaction:

E_{AU} , Uphill- E_A [J]

e.g.: $\Delta U = E_{\text{AU}} - E_{\text{AD}}$ [J]

E_{AD} , Downhill- E_A [J]

Condensation: The phenomenon of going from gaseous state to the liquid state (see physics - matter).

C. Reaction: A reaction in which two smaller molecules combine to form a larger one. Water is invariably one of the products of such a reaction.

Dalton's Law of Partial Pressures: see pressure.

ρ - **Density:** Mass of a substance per unit volume;

$$\rho = m/V \text{ [kg/m}^3\text{]}$$

e.g. Flour: $\rho(\text{F}_2) = M_{\text{F}_2}/V_m = 38/22.4 \text{ [g} \cdot \text{mol}/(\text{l} \cdot \text{mol})] = \text{[g/l]}$

m, mass [kg]

V, volume [m^3]

R, gas constant 8314 [kJ/mol]

Diffusion: The spreading of one substance through another one;

the speed of diffusion can be expressed as:

$$v = \sqrt{(3 \cdot R \cdot T/M)}$$

T, temperature [K]

M, molar mass [g/mol]

Effusion: The process by which a gas under pressure escaped from one compartment of a container to another by passing through a small opening.

Graham's Law of E.: The rate of effusion of a gas is inversely proportional to the square root of its molar mass:

$$t_A/t_B = \sqrt{(M_A/M_B)}$$

t, time [s]

M, molar mass [g/mol]

Gas: A fluid form of matter that fills the container it occupies and can easily be compressed into a much smaller volume (gas is a substance at higher temperature than its critical temperature, vapor is a gaseous form of matter at a temperature below its critical temperature).

1 [mol] ($6.022 \cdot 10^{23}$ molecules) of an ideal gas at $0[^\circ\text{C}]$ and 1[bar] occupies **22.414** [l] = mole-volume.

Ideal G.: A hypothetical gas whose pressure-volume-temperature behavior

can be completely accounted for by the ideal gas equation

• **Ideal G. EQ.:** It expresses the relationships among .

p, V, T, and amount of

$$p \cdot V = n \cdot R \cdot T \quad \text{[kg} \cdot \text{m]} \text{ gas}$$

n, molar amount [mol]

R, gas c. 8,314 [J/(K·mol)]

V, volume [l] [m^3]

T, Temperature [K]

Real G.: Effects of intermolecular forces and the fact that the molecules or atoms involved possess small but distinct volume deviate ($n = P \cdot V/R \cdot T$) versus p significantly;

• **Van der Waals EQ:** An equation that describes

p, V, and T of a non-ideal gas.

$$p_{\text{Ideal}} = p_0 + a \cdot n^2/V^2 \quad \text{[N/m}^2\text{]} = \text{[Pa]}$$

$$n \cdot R \cdot T = (p_0 + a \cdot n^2/V^2) \cdot (V - n \cdot V_X) \quad \text{[kg} \cdot \text{m]}$$

p_0 , observed pressure [Pa]

a, proportionality c. [-]

n, number of molecules [mole]

V, volume [m^3]

V_X , molecular volume [m^3]

Kinetic G. Theory: A theory of properties of an ideal gas in which

point-like molecules with no interactions between them, are in continuous motion at temperature > 0 [K]:

1. A gas is composed of atoms or molecules, which are far apart from each other, empty space in-b/w.
2. Gas molecules are in constant random (brownian) motion along a straight line until collision with the container wall or other gas-molecules; the higher the temperature, the faster they move;
3. The molecules exert no (repelling/attracting) force on each other or on the container; collisions are elastic - their total energy of the molecules remains constant; hence pressure is directly related upon the speed and number of collisions of molecules against the container walls;
4. The average kinetic energy (KE) of the molecules of a gas is proportional to the absolute temperature.

v_D - **Maxwell-Boltzmann Speed Distribution:** Displays the most probable spectrum of molecular speeds available to the system at a particular temperature; e.g.: molecular oxygen has an average speed of 200[m/s] at 73[K], whereas it increases to about 400[m/s] at 273[K]:

$$\begin{aligned} KE &= \frac{1}{2} \cdot m \cdot v_D^2 && [\text{kg} \cdot \text{m}^2 / \text{s}^2] \\ p \cdot V &= n \cdot N_A \cdot KE \cdot 2/3 = n \cdot R \cdot T && [\text{m}^3 \cdot \text{N} / \text{m}^2] \\ &= N_A \cdot KE \cdot 2/3 = R \cdot T && [\text{kg} \cdot \text{m} / \text{mol}] \\ &= N_A \cdot m \cdot v_D^2 \cdot 2/6 = M \cdot v_D^2 \cdot 1/3 \rightarrow v_D = \sqrt{(3 \cdot R \cdot T / M)} \text{ see diffusion.} \end{aligned}$$

N_A , Avogadro c. $6,022 \cdot 10^{23}$ [1/mol]
 v , speed [m/s]
 m , mass [g]
 n , molar amount [mol]
 R , gas c. 8,314 [J/(K·mol)]
 T , temperature [K]
 M , molar mass [g/mol]

Charles and Gay Lussac's Law: The volume of a fixed amount of gas maintained at constant pressure is directly proportional to the absolute temperature of a gas; i.e.:

the volumes of the reactants and the gaseous products are always in the ratio of small whole numbers;

$$V_1/T_1 = \text{constant} = V_2/T_2$$

e.g.: 2·vu of CO(g) + 1·vu of O₂(g) → 2·vu of CO₂(g)

V , volume [m³] or [l]
 T , temperature [K]
 vu , volumetric unit 22.41 [l]

Henry's Law: The solubility of a gas in a liquid is proportional to the pressure of the gas over the solution; see chemistry liquids.

Kinetic Theory (kinetic molecular theory); A theory of properties of an ideal gas in which point-like molecules are in continuous random motion and there are no interactions between them.

Maxwell-Boltzmann Speed Distribution: see gas -kinetic gas theory.

X - Mole Fraction: see pressure, or chemistry stoichiometry - concentration units

Phase: A particular state of matter; a homogeneous part of a system in contact with other parts of the system but separated from them by a well defined boundary; a substance may exist in solid, liquid, and gas phases and in certain cases, in more than one solid phase; e.g.: white and gray tin are two solid phases of tin; ice, liquid, and vapor are three phases of water; etc. see chemistry introduction.

P. Change: Transformation from one phase to another.

P. Diagram: A summary in graphical form of the conditions of temperature and pressure at which the various phases of a substance exist; the cross point of vaporization-, and fusion conduct to the triple point.

Pressure: Force applied per unit area.

Atmospheric P.: The pressure exerted by Earth's atmosphere.

Critical P.: The minimum pressure necessary to bring about liquefaction at the critical temperature.

p_x - **Partial P.:** The pressure of one component in a mixture of gases it would exert if it alone occupied the container:

$$p_{\Sigma} = p_A + p_B + p_C + p_D + \text{etc.}$$

Dalton's Law of partial P.: The partial pressure of a gas in a mixture is independent of other gases present; the total pressure is the sum of the partial pressure of all gasses present:

$$p_A = x_A \cdot p_T \quad [\text{Pa}]$$

Standard Atmospheric P.: The pressure that supports a column of mercury exactly 76 cm high at 0°C [C] at sea level = 101[kPa].

Standard Temperature and P. (STP): 0°C [C] and 101[kPa].

X- Mole Fraction: A dimensionless quantity expressing the ratio of moles of one component to the number of moles of all components present; see stoichiometry - concentration units; e.g.: 2-gas mixture A & B:

$$X_A + X_B = 1$$

$$X_A = n_A / (n_A + n_B) \quad [-]$$

p , pressure [N/m²] [Pa]
 1 bar = 10⁵ [Pa]
 1 atm = 1.01325 [bar]
 1 atm = 101.325 [kPa]

p_x , partial pressure [N/m²] [Pa]

x_A , mole fraction [-]
 p_T , total pressure [Pa]

n , molar amount [mol]

Radical: see stoichiometry.

Sublimation: The process in which molecules go directly from the solid into the vapor phase.

Temperature: How hot or cold a sample is; the intensive property that determines the direction in which heat will flow between two objects in contact.

Critical T.: The temperature above which a substance cannot exist as a liquid i.e.: gaseous.

Triple Point: The point at which the vapor, liquid, and solid states of a substance meet in a phase diagram and are in dynamic equilibrium; e.g.: TP of H₂O = 0.0099°C [C] and 0.61[kPa]

Van der Waals Equation: see gas - real gases.

Vapor: The gaseous phase of a substance (specifically, of a substance that is a liquid or a solid at the temperature in question); see chemistry - liquid.

V. Pressure: The pressure exerted by the vapor of a liquid (or solid) when the vapor and the liquid (or solid) are in dynamic equilibrium; the higher the intermolecular forces the lower the vapor pressure; e.g.: H₂O(g) ↔ H₂O(l)

- **VP Lowering:** see colligative properties.

Glossary - Chemistry (Liquid)

Boiling Point: The temperature at which the vapor pressure of a liquid is equal to the external atmospheric pressure ($KE_{\text{Liquid}} = KE$ of surrounding environment); the higher the intermolecular forces the lower the vapor pressure the higher is also the boiling point

BP Elevation: see colligative properties.

Colligative Properties: A property that depends only on the number of solute particles present in a solution and not on their chemical composition;

Boiling Point Elevation: The vapor pressure at any temperature of a given solution is lower than that of the pure solvent; the liquid- vapor curve for the solution lies below that of the pure solvent intersecting the boiling point at a higher temperature: $\Delta T_b = k_{\text{Slt}} \cdot b_{\text{Sle}}$ [°C]

ΔT_s , change in temperature [°C]
 b_{Sle} , molality of solute [mol/g]
 $k_{\text{b-Slt}}$, boiling point coefficient of solvent [°C/(g·mol)]

Freezing Point Depression: The lowering of the freezingpoint of a solution caused by the presence of a solute; when the solution begins to freeze, only the solvent solidifies and solute is left behind in the solution, raising the concentration of the solute as well as lowering the freezing point of the solvent:

ΔT_s , change in temperature [°C]
 $k_{\text{f-Slt}}$, freezing point coefficient of solvent [°C/(g·mol)]

$\Delta T_f = k_{\text{Slt}} \cdot b_{\text{Sle}}$ [°C]

Osmosis: (Gk. osmos, impulse or thrust) The diffusion of water, or any other solvent, across a differentially permeable membrane; in the absence of other forces, the movement of water during osmosis will always be from the region of greater potential to one of lesser water potential.

- **O. Pressure:** The pressure needed to balance the flow of solvent through a semipermeable membrane; i.e.: the pressure required to stop osmosis:
 $\pi = c \cdot R \cdot T = n \cdot R \cdot T / V$ [mol·N·m·K/(K·mol·m³)] = [N/m²] = [Pa]
- **Reverse O.:** The passage of solvent out of a solution when a pressure greater than the osmotic pressure is applied on the solution side of a semipermeable membrane; e.g.: a method of desalination using high pressure to force water through a semipermeable membrane from a more concentrated solution to a less concentrated solution.

c , molar concentration [mol/l]
 R , gas constant 8,314 [J/K·mol]
 T , absolute temperature [K]
 n , molar amount [mol]
 V , volume [m³]

Vapor Pressure Lowering: The presence of a non-volatile solute raises the boiling point of the solvent, because it can block the escape of solvent molecules but has no effect on the rate of return of the solvent particles from the vapor to the solution.

Complex Formation: A metal ligand coordinate-covalent bond formation.

Complex Ion: Ions containing a central metal cation bonded to one or more molecules or ions like the following sample show: $\text{Ag}(\text{NH}_3)_2^+$, CdCl_4^{2-} , $\text{Cu}(\text{NH}_3)_4^+$, $\text{Fe}(\text{CN})_6^{3-}$, $\text{Fe}(\text{H}_2\text{O})_6^{3+}$, PtCl_4^- , etc.

e.g.: $\text{AgCl}(\text{s}) + 2\text{NH}_3(\text{aq}) \leftrightarrow \text{Ag}(\text{NH}_3)_2^+(\text{aq}) + \text{Cl}^-(\text{aq})$

according to Le Chatelier's principle, the removal of Ag^+ ions from the solution to form $\text{Ag}(\text{NH}_3)_2^+$ ions will cause more AgCl to dissolve, whereas if NH_3 is absent, more AgCl would precipitate.

Ligand: A group attached to a central metal ion in a complex, i.e.: molecules or ions that surround the metal atom and the ligands can be thought of a Lewis acid-base reaction; every ligand has at least one unshared pair of valence electrons (lone pairs), hence is a Lewis base; e.g.: H_2O , NH_3 , CO , Cl^- , etc.

K_f CF. Constant: A measure of the tendency of a metal to form a particular complex ion; the larger K_f , the more stable the complex ion;

Concentration Units: see chemistry stoichiometry.

Condensation: The phenomenon of going from gaseous state to the liquid state (see physics - matter).

C. Reaction: A reaction in which two smaller molecules combine to form a larger one. Water is invariably one of the products of such a reaction.

Dalton's Law of Partial Pressures: The partial pressure of a gas in a mixture is independent of other gases present; the total vapor pressure is the sum of the partial pressure of all gasses present (see there):

Deposition: The process in which the molecules go directly from the vapor into the solid phase.

Desalination: Purification of seawater by the removal of dissolved salts; see osmosis.

Diffusion: (L. diffundere, to pour out) The net movement of suspended or dissolved particles from a more concentrated region to a less concentrated region by virtue of their kinetic energy (result of the random movement, brownian motion - see physics-matter) of individual molecules; the process tends to distribute such particles uniformly throughout the medium; see also chemistry - gas.

Dilution: see solution.

Distillation: The separation of a mixture by making use of the different volatilities of its components;

Fractional D.: A procedure for separating liquid components of a solution that is based on their different boiling points; based on Dalton's law of partial pressures;

see chemistry-gas:

$$p_A = x_A \cdot p \quad [\text{Pa}]$$

x_A , mole fraction [-]

p_T total pressure [Pa]

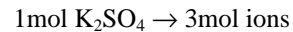
Vacuum D.: Boiling off humidity by lowering the temperature below freezing and subsequently decreasing the pressure; e.g.: freeze-dried coffee.

Electrolyte: 1) An ionically conducting medium. 2) A substance that, when dissolved in water, results in a solution that can conduct electricity; see table below.

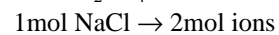
E. Rule: For a net potential of zero, the positive and negative charges must add up to zero; a solution must contain essentially as many anionic as cationic charges.

Non-E.: Is a solution in which no proportion of the solute molecules are ionized, hence does not conduct electricity; e.g.: $\text{C}_6\text{H}_{12}\text{O}_6(\text{aq})$; see table below

Strong E.: Is a solution in which a large proportion of the solute molecules are ionized (complete dissociation into ions);



e.g.: $\text{NaCl}(\text{s}) \rightarrow \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq})$



Weak E.: Is a solution in which only a small proportion of the solute molecules are ionized (partly dissociation into ions); e.g.: $\text{CH}_3\text{COOH}(\text{aq})$; see table below

Equilibrium: The state of final balance of a multi-compound homogenous mixture;

for K_C , Q_C , Dynamic E. etc., see chemistry-stoichiometry equilibrium.

Evaporation: The escape of molecules from the surface of a liquid; also called vaporization.

Freezing Point: The temperature at which a liquid freezes (crystallization); the normal freezing point is the freezing temperature under a pressure of 1 atm, i.e.: liquid and solid phase are in equilibrium.

FP Depression: see colligative properties.

Henry's Law: The solubility of a gas in a liquid is proportional to the pressure of the gas above the solution; solubility $\propto c$; applicable only if: solution is low in concentration;

low pressures; solvent does not react with gas.

k , Henry's variable [mol/(l·Pa)]

$$c = k \cdot p \quad [\text{mol/l}]$$

p , partial pressure [1/atm] [Pa]

Hydration: A process in which an ion or a molecule is surrounded by water molecules arranged in a specific manner; e.g.: water - H_2O molecules attach to a central ion (ion-dipole interaction).

Hydrated Anion: Hydrogen bonds form between the H of water and the central anion; e.g.: SO_4^{2-} .

Hydrated Cation: Ion-dipole forces between the O of water and the central ion; e.g.: Be^{2+} .

H. Crystals: Hydrated ions remain intact even in a solidified structure;

e.g.: $[\text{Fe}(\text{OH}_2)_6]^{3+} \text{Cl}_3^{3-}$ actual structure

notation: $\text{Fe}_2\text{Cl}_3 \cdot 6\text{H}_2\text{O}$

or $[\text{Cu}(\text{OH}_2)_4]^{2+} [\text{SO}_4(\text{H}_2\text{O})]^{2-}$

notation: $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$

Enthalpy of H.: see chemistry - thermochemistry.

Hydrophilic: Water-liking.

Hydrophobic: Water-fearing.

Intermolecular Forces: see chemistry molecule.

Law of Mass Action: see chemistry stoichiometry - equilibrium

Ligand: see complex formation.

Liquid: A fluid form of matter that takes the shape of the part of a container it occupies and is almost incompressible.

Melting Point: The temperature at which solid and liquid phases coexist in equilibrium.

Miscible: Two liquids that are completely soluble in each other in all proportions are said to be miscible.

Mixture: A combination of two or more substances in which the substances retain their identity.

Heterogeneous M.: The individual components of a mixture remain physically separated and can be seen as separate components.

Homogeneous M.: The composition of the mixture, after sufficient stirring, is the same throughout the solution.

Osmosis: see colligative properties.

Precipitation Reaction: see solution-properties.

Phase: see chemistry introduction.

Raoult's Law: The vapor pressure of an ideal solution of a nonvolatile solute is directly proportional to the mole fraction of the solvent in the solution; i.e.: The partial pressure of the solvent over a solution is given by the product of the vapor pressure of the pure solvent and the mole fraction of the solvent in the solution.

$$p_{\text{Slt}} = x_{\text{Slt}} \cdot p_{\text{Slt}}^{\circ} \quad [\text{Pa}]$$

$$p_{(\Sigma)} = p_{\text{Slt}} + p_{\text{Sle}} + \text{etc.}$$

Deviation of RL.: Non-ideal (real) solutions deviate from the ideal pattern given by RL:

- Negative DoRL.: Total- and partial pressure are less than predicted by RL; attractive forces between solute/solvent is greater than with like ones; the heat of solution is negative (exothermic).
- Positive DoRL.: Total- and partial pressure are greater than predicted by RL; attractive forces between solute/solvent is lesser than with like ones; the heat of solution is positive (endothermic).

Semipermeable Membrane: A membrane that allows solvent molecules to pass through, but blocks the movement of solute molecules; see osmosis.

Solubility: The maximum amount of solute that can be dissolved in a given quantity of a specific solvent at a specific temperature (for gases: at a specific pressure); the concentration of a saturated solution of a substance; e.g.: how much of a salt can be dissolved in a solvent.

K_S - **S. Constant:** see solubility product;

K_{sp} - **S. Product:** The product of relative ionic molar concentrations of the constituent ions in a saturated solution, each raised to the power of its stoichiometric coefficient in the equilibrium EQ;

$$K_{sp} = K_C \cdot c_{(\text{HA})} = c_{(\text{A}^-)} \cdot c_{(\text{C}^+)}$$

e.g.: $\text{Hg}_2\text{Cl}_2(\text{s}) \leftrightarrow \text{Hg}_2^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq})$;

K_{sp} - **S. Quotient:** The molar analogue of the solubility product, but with the molar concentrations not necessarily those at equilibrium.

$Q_{sp} \geq K_{sp}$ precipitate will form, whereas if $Q_{sp} < K_{sp}$, still more salt can be added and will dissolve.

S. Rules: Solubility pattern of a range of common compounds in water - see table below.

- unpolar and polar substances are not miscible; e.g.: oil and water.
- like dissolves like; e.g.: ionic bonded element dissolve well in polar solvents, NaCl in H_2O .

Pressure and Solubility: The process always implies a reduction of volume:

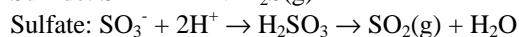
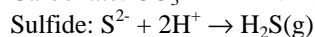
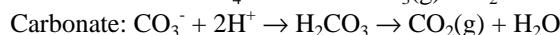
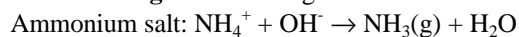
S. in **Liquids:** Since liquids are almost incompressible there is little influence in a change of volume.

S. in **Gases:** With increasing pressure more gas becomes dissolved in the solution; e.g.: soft-drink bottle effect, diving (see Henry's law).

Properties of aqueous solutions:

Acid-Base Reaction: Is a reaction in which protons are transferred; it is a proton transfer reactions; see chemistry acid-base.

Gas-Forming Reaction: e.g.:



Oxidation-Reduction Reaction (RedOx): A reaction in which electrons are transferred; it is an electron transfer reaction; see chemistry - stoichiometry-redox.

Precipitation Reaction: A reaction in which an insoluble solid product (which separates from the solution) is formed when two solutions are mixed; i.e.: $\text{AgNO}_3(\text{aq}) + \text{NaCl}(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{NaNO}_3(\text{aq})$

Temperature and Solubility: Temperature always influences solubility in a negative or positive way:

Endothermic S.: With increasing temperature, solubility increases, the endothermic (energy consuming) process can proceed; follows the principle of Le Chatelier (see chemistry stoichiometry)

Exothermic S.: Solubility decreases with lower temperatures; gases are exothermic (energy releasing) even though solubility decreases with raising temperatures.

Sln - Solution: A homogeneous mixture of two or more substances.

Aqueous S.: A solution in which the solvent is water.

Enthalpy of S.: see chemistry - thermochemistry.

S. Concentration: The amount of solute present in a given quantity of solution.

Ideal S.: Any solution that obeys Raoult's law at any concentrations. Real solutions resembles ideal solutions more closely the lower the concentration; below 0.1[mol/kg] for non electrolytic solutions and 0.01[mol/kg] for electrolytic solutions.

Dilution: A procedure for preparing a less concentrated solution from a more concentrated solution;

Reminder: $n_{(\text{concentrated solvent})} = n_{(\text{diluted solvent})}$

If a solute is added in very small quantities compared to the solvent, than the vapor pressure of the liquid can be said to be equal to that of the pure solvent; any diluted liquid given as a %-value usually refers to mass-% in a 100g of solution;

being so diluted Raoult's law can be implemented: Δp , change of vapor pressure of solution

molar fraction of solute B = $x_{(B)} = \Delta p / p^{\circ}_{(A)}$ [-] $p^{\circ}_{(A)}$, vapor pressure of pure solvent [Pa]

Saturated S.: At a given temperature, the solution that results when the maximum amount of a substance has dissolved in a solvent; dissolved and undissolved solute are in dynamic equilibrium.

- **Oversaturated S.:** (supersaturated) A solution that contains more solute than it has the capacity to dissolve (unstable).

- **Unsaturated S.:** A solution that contains less solute than it has the capacity to dissolve.

Standard S.: A solution of accurately known concentration, used for acid-base titrations to calculate the dosage of molar amounts out of a used volume; see chemistry-acid-base.

Slt - Solvent: The substance (usually one, or more) present in larger amount in a solution.

Sle - Solute: The substance present in smaller amount in a solution.

γ - **Surface Tension:** A measure of the force that must be applied to surface molecules so that they experience the same force as molecules in the interior of the liquid; the amount of energy required to stretch or increase the surface by unit area; decreases with increasing temperature; work has to be expended to bring molecules to the surface.

Titration: see chemistry acid-base.

Triple Point: The point at which the vapor, liquid, and solid states of a substance are in equilibrium.

Vapor: The gaseous phase of a substance (specifically, of a substance that is a liquid or a solid at the temperature in question).

V. Pressure: The pressure exerted by the vapor of a liquid (or solid) when the vapor and the liquid (or solid) are in dynamic equilibrium; the higher the intermolecular forces the lower the vapor pressure a liquid's vapor pressure in an open beaker cannot increase since surrounding atmospheric pressure is more or less constant, therefore temperature (at boiling point) remains constant until the entire liquid has evaporated;

e.g.: $\text{H}_2\text{O}(\text{g}) \leftrightarrow \text{H}_2\text{O}(\text{l})$

- **VP Lowering:** see colligative properties.

Enthalpy of V.: see thermochemistry.

Vaporization (evaporation): The escape of molecules from the surface of a liquid; a certain number of molecules in a liquid (at any temperature) possess sufficient energy to escape from the surface; see chemistry gas - Maxwell-Boltzman speed distribution.

Molar Heat of Vaporization: see chemistry - thermochemistry;

Viscosity: A measure of a fluid's resistance to flow; decreases with increasing temperature.

Volatility: The readiness with which a substance vaporizes: A substance is typically regarded as volatile if its boiling point is below 100°C.

Volatile: as a measurable vapor pressure.

Water: Liquid substance of H-O-H molecules, in which intermolecular forces (H-bonds) are responsible for the high boiling point; water molecule is bent due to the lone pair of the O and possess a dipole moment, since the O-atom (more electronegative) deprives the H-atoms off their electron clouds.

Hard W.: Water that contains Ca^{2+} and Mg^{2+} ions.

Soft W.: Water that mostly free of Ca^{2+} and Mg^{2+} ions.

Molal boiling point elevation and freezing point depression constants of several common compounds:

Solvent	freezing point [°C]	$k_{f, \text{Slt}}$ [°C·kg/mol]	boiling point [°C]	$k_{b, \text{Slt}}$ [°C·kg/mol]
Acetic acid	16.6	-3.90	117.9	2.93
Benzene	5.5	-5.12	80.1	2.53
Carbon Tetrachloride	-22.8	-29.8	76.8	5.02
Chloroform	-63.5	-4.68	61.2	3.63
Cyclohexan	6.6	-20.0	80.7	2.79
Ethanol	-117.3	-1.99	78.4	1.22
Naphtalin	80.2	-6,8	-	-
Water	0	-1.86	100	0.52

Classification of solutes in aqueous solution

Strong Electrolyte	Weak Electrolyte	Non- Electrolyte
HCL	CH ₃ COOH	(NH ₂) ₂ CO (urea)
HNO ₃	HF	CH ₃ OH (methanol)
HClO ₄	HNO ₂	C ₂ H ₅ OH (glucose)
H ₂ SO ₄	NH ₃	C ₆ H ₁₂ O ₆ (glucose)
NaOH	H ₂ O	C ₁₂ H ₂₂ O ₁₁ (sucrose)
Ba(OH) ₂		
ionic compounds		

Solubility Rule (based on Group-I &-II &-III elements, NH₄⁺ Al³⁺ Pb²⁺ Sn²⁺):

	soluble compounds	semi-soluble	insoluble compounds
Bromides (Br ⁻)	all*	Pb ²⁺ , Hg ²⁺	Ag ⁺ , Hg ₂ ²⁺ , TI ⁺
Carbonates (CO ₃ ²⁻)	Group-I elements, NH ₄ ⁺		all remaining*
Chlorides (Cl ⁻)	all*	Pb ²⁺	Ag ⁺ , Hg ₂ ²⁺ , TI ⁺
Chlorates (ClO ₃ ⁻)	all*		
Fluorides (F ⁻)	single charged cations		multiple charged cations
Hydroxides (OH ⁻)	Group-I elements Ba ²⁺ , NH ₄ ⁺	Ba ²⁺ , Ca ²⁺ , Sr ²⁺ Mg ²⁺	all remaining
Nitrates (NO ₃ ⁻)	all*		
Sulfates (SO ₄ ²⁻),	all*	Ca ²⁺ , Ag ⁺	Sr ²⁺ Ba ²⁺ , Pb ²⁺ , Hg ₂ ²⁺
Sulfides (S ²⁻)	Group-I & -II elements, NH ₄ ⁺ ,		all remaining; (Fe ₂ ³⁺ , Al ³⁺ , Cr ³⁺ form hydroxides)
Sulfites (SO ₃ ²⁻)	Group-I elements, NH ₄ ⁺ ,		all remaining*
Oxides (O ²⁻)	all Group-I elements, NH ₄ ⁺ nitrate (NO ₃ ⁻), chlorate (ClO ₃ ⁻), perchlorate (ClO ₄ ⁻)		most hydroxides (OH ⁻)
Phosphates (PO ₄ ³⁻)	Group-I elements, NH ₄ ⁺		all remaining*

*) elements taken in consideration: Group-I & -II elements, NH₄⁺, Ag⁺, TI⁺, NH₄⁺, Mn²⁺, Fe²⁺, Co²⁺, Ni²⁺, Cu²⁺, Zn²⁺, Cd²⁺, Hg₂²⁺, Sn²⁺, Pb²⁺, Al³⁺, Cr³⁺, Fe³⁺

Compounds of group-I elements: H⁺, Li⁺, Na⁺, K⁺, etc.

Compounds of group-II elements: Be²⁺, Mg²⁺, Ca²⁺, Sr²⁺, etc.

Compounds of group-III elements: transition metals like Fe, V, Au, etc.

Glossary - Chemistry (Solid)

Amorphous solid: A solid that lacks a regular 3-dimensional arrangement of atoms or molecules;
i.e.: a liquid with an extremely high viscosity index e.g.: glass, tar, etc.

Bond: see chemistry molecule.

Born-Haber Cycle: The cycle that relates lattice energies of ionic compounds to ionization energies, electron affinities, heats of sublimation and formation, and bond dissociation energies; see thermochemistry.

Closest Packing: The most efficient arrangements for packing atoms, molecules, or ions in a crystal.

Cohesion: The intermolecular attraction between like molecules (see physics - matter).

Colligative Properties: see chemistry - liquid.

Conductor: Substance capable of conducting electric current (see physics - electromagnetism).

Coordination Number: see lattice.

Crystal: An regular arrangement of atoms, ions, and molecules of periodically repeated, identically constituted, congruent lattice consisting of unit cells; see table below.

Crystalline Solid: A solid that possesses rigid and long-range order; its atoms, molecules, or ions occupy specific positions; e.g.: NaCl, diamond, graphite, etc.

Fractional C.: The separation of a mixture of substances into pure components on the basis of their differing solubility.

Crystallization: The process in which dissolved solute comes out of solution and forms crystals.

Recrystallization: Purification by repeated dissolving and crystallization.

Types of crystals: see table below.

Covalent C.: The structure is primarily determined by the geometry of the bonds formed by each atom; the lattice ranges from face-centered cubic (diamond) to cubic close-packed (tetrahedral of SiO₄ form SiO₂) and to triangular pyramidal (As₄).

Ionic C.: Are composed of charged species with anions and cations different in size; the lattice is face-centered (NaCl and most others $c_N = 6$) or simple cubic (CsCl, CsBr, CsI, $c_N = 8$) where every anion tends to gather as many cations and vice versa (1:1 structure).

Metallic C.: Generally body-centered, face-centered, or hexagonal close-packed with the bonding electrons delocalized over the entire crystal; the cohesive of these e-forces are responsible for the metals strength;

All metals, except Mn, Hg, Ga, In, Ge, Sn have one of the 3 types of structures:

- **Monoatomic Body-Centered:** The lattice is body-centered cubic, one atom is situated at each lattice point, giving 2 atoms per unit cell ($1 + 8 \cdot \frac{1}{8}$); $c_N = 8$; e.g.: Li, Na, K, V, Cr, Fe, Rb, Nb, Mo, Cs, Ba, Ta, W
- **Cubic Close-Packed:** The lattice is face-centered cubic, one atom situated at each lattice point, giving 4 atoms per unit cell ($6 \cdot \frac{1}{2} + 8 \cdot \frac{1}{8}$); $c_N = 4$; e.g.: Al, Ca, Ni, Cu, Sr, Rh, Pd, Ag, IR, Pt, Au, Pb
- **Hexagonal close-packed:** The close-packed layers are perpendicular to the vertical axis or perpendicular to the body's diagonal of the unit cell; e.g.: Be, Mg, Sc, Ti, Co, Zn, Y, Zr, Tc, Ru, Cd, La, Hf, Re, Os, Ti

Molecular C.: Structure is determined by the size and the shape of the atoms involved and range from face-centered cubic (CO₂) to face-centered orthorhombic (Cl₂, Br₂, I₂)

Deposition: The process in which the molecules go directly from the vapor into the solid phase.

Freezing Point: The temperature at which a liquid freezes (crystallization); the normal freezing point is the freezing temperature under a pressure of 1atm, i.e.: liquid and solid phase are in equilibrium;

Lattice: A regular periodic framework throughout an area or space, made of interwoven strips, as in crystals.

L. **Energy:** The energy required to completely separate one mole of a solid ionic compound into gaseous state; LE depends upon the charge of ions (the higher, the stronger), and the size of ions (the smaller, the stronger the force of attraction).e.g.: $\text{Na}^+(\text{g}) + \text{Cl}^-(\text{g}) \rightarrow \text{NaCl}(\text{s}) \quad \Delta H_L = -788 \text{ [kJ/mol]}$;
i.e.: the energy needed for vaporizing this solid +788 [kJ/mol]; see thermochemistry.

L. **Enthalpy:** The standard enthalpy change for the conversion of an ionic solid to a gas of ions.

c_N - L. **Coordination Number:** The number of nearest neighbors of an atom (ion), in a solid (of opposite charge); for complex ions (see chemistry stoichiometry) it is the number of atoms directly attached to the central metal ion; the higher c_N , the stronger the intermolecular bond; see unit cell.

L. **Points:** the positions occupied by atoms, molecules, or ions that define the geometry of a unit cell.

Melting Point: The temperature at which solid and liquid phases coexist in equilibrium.

Metals: 1) A substance that conducts electricity as well as heat, has a metallic luster, is malleable and ductile, forms cations and has basic oxides. 2) A metal consists of cations held together by a sea of electrons (have the tendency to form positive ions in ionic compounds); i.e.: iron, copper, uranium, etc.

Phase: see chemistry - introduction.

Rays: (see physics - nuclear).

α **Alpha R.:** Helium ions with a positive charge of +2; also known as alpha particles.

β **Beta R.:** Electrons.

γ **Gamma R.:** High energy radiation.

X-Ray: Electromagnetic radiation with wavelengths ranging from 10[pm] to about 1000[pm]; see physics - optics.

XR-Diffraction: The analysis of crystal structures by studying the interference pattern in a beam of x-rays:	θ , diffraction angle	[°]
To determine the bond length:	λ wavelength	[m]
$2 \cdot d \cdot \sin\theta = n \cdot \lambda$	n, integer, 1, 2, 3,	[-]

Salt: An ionic compound made up of a cation other than H^+ and an anion other than OH^- or O^{2-} .

S. **Hydrolysis:** The reaction of the anion or cation, or both, of a salt with water.

Solid: A rigid form of matter that maintains the same shape whatever the shape of the container.

Sublimation: The process in which molecules go directly from the solid into the vapor phase; can be measured easily in a vacuum chamber where at a given temperature a distinct vapor pressure can be detected (solid and gas phases are in equilibrium).

S. **Vapor Pressure:** The pressure exerted by the vapor of a solid when the vapor and or solid are in dynamic equilibrium; the higher the intermolecular forces the lower the vapor pressure; see chemistry - liquid.

Triple Point: The point at which the vapor, liquid, and solid states of a substance are in equilibrium; see chemistry - gas.

Unit Cell: The basic repeating unit of the arrangement of atoms, molecules, or ions in a crystalline solid; when stacked together can reproduce the entire crystal; see table below.

Sc - **Simple Cubic** (primitive cubic): The unit cell of a space lattice is obtained by joining 8 points to give a solid with pairs of parallel faces (cube); $c_N = 6$.

Bcc - **Body-Centered Cubic:** A cube with an additional point in the center in the middle of the cube; $c_N = 8$.

Fcc - **Face-Centered Cubic:** A cube with a n additional point in the center of each face; $c_N = 12$.

Monoclinic: $a \neq b \neq c$; $\alpha = \gamma = 90^\circ$ ($\beta \neq 90^\circ$)

Hexagonal: $a = b \neq c$; $\alpha = \beta = 90^\circ$ ($\gamma = 120^\circ$)

Orthorhombic: $a \neq b \neq c$; $\alpha = \beta = \gamma = 90^\circ$

Rhombohedral: $a = b = c$; $\alpha = \beta = \gamma \neq 90^\circ$

Tetragonal: $a = b \neq c$; $\alpha = \beta = \gamma = 90^\circ$

Triclinic: $a \neq b \neq c$; $\alpha \neq \beta \neq \gamma \neq 90^\circ$

Common lattice energies

Type of compound	Compound	Ions	Σ of ionic radius	Lattice E. [kJ/mol]
1+, 1-	NaCl	Na^+, Cl^-	$95 + 181 = 276$	- 788
	CsCl	Cs^+, Cl^-	$169 + 181 = 350$	- 669
1+, 2-	Na_2O	$2\text{Na}^+, \text{O}^{2-}$	$95 + 140 = 235$	-2570
	Cs_2O	$2\text{Cs}^+, \text{O}^{2-}$	$169 + 140 = 309$	-2090
2+, 1-	MgCl_2	$\text{Mg}^{2+}, 2\text{Cl}^-$	$65 + 181 = 246$	-2525
2+, 2-	MgO	$\text{Mg}^{2+}, \text{O}^{2-}$	$65 + 140 = 205$	-3890

The cubic unit cell

Type of cell	# of points at corners	# of points in faces	# of points in center of cube	total	c_N
scc	$8 \times \frac{1}{8}$	0	0	1	6
bcc	$8 \times \frac{1}{8}$	0	1	2	8
fcc	$8 \times \frac{1}{8}$	$6 \times \frac{1}{2}$	0	4	12

Types of crystals and general properties

Type of crystal	Particles involved	Force(s) holding the unit together	General properties	Examples
Ionic	positively and negatively charged ions	electrostatic attraction	hard, brittle, high melting point, poor conductor of heat and electricity	NaCl, LiF, MgO, CaCO_3
Covalent (spatial) 3-dimensional	atoms	covalent bond	hard, high melting point, poor conductor of heat and electricity	C (diamond), SiO_2 (quartz)
Covalent (layers) 2-dimensional	atoms	covalent bond + dispersion forces	soft, high melting point, good conductor of electricity	C (graphite)
	atoms and ions	covalent bond + electrostatic attract.	high melting point, poor conductor of electricity	mica, clay
Covalent (chain) 1-dimensional	atoms	covalent bond + dispersion or dipole-dipole forces	fibrous, partly meltable to a viscous liquid,	SiS_2
	atoms and ions	covalent bonds + electrostatic attraction	fibrous, poor conductor of heat and electricity	asbestos
Molecular (polar)	polar molecules	dispersion-, dipole-dipole-forces, hydrogen bonds	soft, low melting point, poor conductor of heat and electricity	Ar, CO_2 , I_2 , H_2O , $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ (sucrose)
	nonpolar molecules	dispersion forces only		H_2 , Cl_2 , CH_4
Metallic	positive ions (cations) and highly motile electrons	metallic bond	soft to hard, low to high melting point, good conductor of heat and electricity, easily malleable	all metallic elements; Na, Mg, Fe, Cu, etc.

Glossary - Chemistry (Thermochemistry)

Adiabatic Process: A process in which no heat exchange occurs between the system and its surrounding environment (see physics - heat).

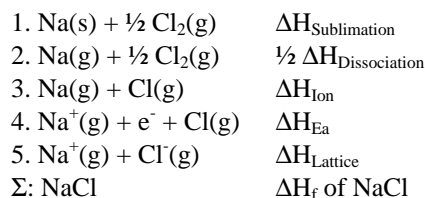
Bond: A link between bonds - see chemistry - molecule.

Double B.: Two electron pairs shared by neighboring atoms.

Multiple B.: Bonds formed when two atoms share two or more pairs of electrons.

Triple B.: Three electron pairs shared by two neighboring atoms.

Born-Haber Cycle: The cycle that relates lattice energies of ionic compounds to ionization, energies electron affinities, heats of sublimation, and formation, and bond dissociation energies; e.g.: ΔH_f of NaCl is obtained out of the energy differences of steps 1 to 5



Calorimeter: A device to measure the heat released/absorbed by a process under constant volume (isochorous) i.e.: $W = 0 \rightarrow \Delta U = Q$; the internal energy of a process will be converted into a change of heat.

Bomb C.: A combustion chamber w/n an isolated, sealed tank with stirrer, igniter and thermometer;

Calorimetry: The use of a calorimeter to measure the thermochemical properties of reactions.

Clausius Clapeyron's Equation: The quantitative relationship between

the vapor pressure p_v of a liquid and the absolute temperature T:

$$\ln(p_v) = C - \Delta H_v / (R \cdot T)$$

Conversion of p_1, T_1 to p_2, T_2 :

$$\log(p_1/p_2) = \Delta H_v / R \cdot \{T_2 - T_1 / (T_2 \cdot T_1)\}$$

p_v , vapor pressure [N/m ²]	[Pa]
C, constant	[1/K]
R, gas constant	8314 [J/mol]
T, temperature	[K]

Collision Theory: The theory of elementary gas-phase bimolecular reactions in which it is assumed that molecules react only if they collide with at least enough kinetic energy for bonds to be broken.

Activated Complex T.: A theory of reaction rates in which it is postulated that the reactants form an activated complex; i.e.: compounds temporarily formed by the reactant molecules as a result of collision.

Activated Complex: A combination of the two reactant molecules that can either go on to form products or fall apart into the unchanged reactants;

e.g.: $A_2 + X_2 \leftrightarrow A_2X_2 \rightarrow 2AX$; where A_2X_2 is very short-living.

E - Energy: The capacity to do work or supply heat; $KE + PE = \text{constant!}$ Units in: $[kg \cdot m^2/s^2] = [N \cdot m] = [J]$

E_A - **Activation Energy:** The minimum energy needed for reaction; i.e.: the height of the activation barrier; usually endothermic, followed by a more or less exothermic reaction; resulting in a net exo-/endothermic reaction (see stoichiometry - kinetics Arrhenius behavior):

Uphill- E_A

[J/mol]

e.g.: $\Delta H = E_{AU} - E_{AD}$ [J/mol]

E_{AD} , Downhill- E_A [J/mol]

Bond E.: The energy needed in diatomic molecules to dissociate 1mol of molecules into gaseous atoms;

e.g.: $H_2(g) \rightarrow 2H(g, \text{atomic}) = 436[kJ/mol]$; i.e.: $H = 218.0[kJ/mol]$

Dissociation E.: The amount of energy needed to break a bond - see table below.

Free E.: Energy available to do work.

U - Internal E. (energy of reaction): The total energy of a system; the sum of KE and PE of all particles; absolute amount of U is not detectable, but relative changes can be expressed as follows:

$$\Delta U = Q + W \text{ (energy supplied as heat and work) [J]}$$

$\Delta U < 0$: internal energy decreases by giving of energy to its surrounding environment;

$\Delta U > 0$: internal energy increases by taking up energy from the surrounding environment;

$\Delta U = 0$: no change in internal energy ($W = 0$; $\rightarrow \Delta U = Q_v$); Q_v = heat supplied at constant volume;

Ionization E.: Energy required to remove of electrons from an element; see ΔH - electron enthalpy.

Kinetic E.: Energy available because of the motion of an object; see physics - mechanics: $KE_L = \frac{1}{2} \cdot m \cdot v^2$ [J]

Lattice E.: The energy required to completely separate one mole of a solid ionic compound into gaseous state; i.e.: energy of + energy is needed to vaporized this solid substance. LE depends upon charge of ions (the higher, the stronger) and the size of ions (the smaller, the stronger the force of attraction):

e.g.: $Na^+(g) + Cl^-(g) \rightarrow NaCl(s)$

$\Delta H_L = -788$ [kJ/mol];

Law of Conservation of E.: see 1st law of thermodynamics.

Potential E.: Energy available by virtue of an object's position; see physics - mechanics: $PE = m \cdot g \cdot y$ [J]

Standard Free E. of Change (ΔG°): The free-energy change

when reactants in their standard states are converted to products in their standard states.

Standard Free E. of Formation (ΔG_f°): The free-energy change when 1 mole of a compound is synthesized from its elements in their standard states.

Thermal E.: Energy associated with the random motion of atoms and molecules.

H - Enthalpy: A thermodynamic quantity used to describe heat changes taking place at constant pressure (isobar); i.e.: reservoir of energy that can be obtained as heat;

$\Delta H = \Delta U + W = H_{\text{final}} - H_{\text{initial}}$	[kJ/mol]	ΔU , internal energy	[J]
$\Delta H = \Delta U + p \cdot \Delta V = \Delta U + R \cdot \Delta n \cdot T$	[kJ/mol]	W, work [N·m]	[J]
$\Delta H < 0$: heat released (exothermic reaction);		p, pressure [N/m ²]	[Pa]
$\Delta H > 0$: heat absorbed (endothermic reaction);		ΔV , change in volume	[m ³]
Reminder: when using ΔH , don't forget to add the reactants- and products phase!		R, gas c. 8.314	[J/(K·mol)]
		Δn , molar amount	[mol]
		T, temperature	[K]

E. of **Chemical Change:** Processes involved in chemical changes;

- E. of **Formation:** see Born Haber cycle.
- E. of **Hydration:** Energy which is released from ions(g) to dissolved hydrated ions(aq) in a hypothetical process; ΔH_H is always negative.
e.g.: $K^+(g) + Cl^-(g) \rightarrow (H_2O) \rightarrow K^+(aq) + Cl^-(aq)$ $\Delta H_H = -684.1$ [kJ/mol]
- E. of **Reaction:** The difference between the enthalpies of the products and the enthalpies of the reactants; measured in [J/mole] (compare physics - heat).
Endothermic R.: Processes that absorb heat from the surrounding environment, $\Delta H > 0$;
i.e.: $\frac{1}{2} H_2(g) + \frac{1}{2} I_2(s) \rightarrow HI(g)$ $\Delta H = +25.9$ [kJ]
Exothermic R.: Processes that give off heat to the surroundings, $\Delta H < 0$;
i.e.: $H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(g)$ $\Delta H = -241.8$ [kJ]
- E. of **Solution:** The heat generated or absorbed when a certain amount of solute is dissolved in a certain amount of solvent; i.e.: the sum of energy required to break the lattice structure and energy set free during solvatization; e.g.: $\Delta H = H_{Sln} - H_{Components}$
 1. $KCl(s) \rightarrow (H_2O) \rightarrow K^+(g) + Cl^-(g)$...lattice energy (positive) $\Delta H_S = +701.2$ [kJ/mol]
 2. $K^+(g) + Cl^-(g) \rightarrow (H_2O) \rightarrow K^+(aq) + Cl^-(aq)$...hydration energy $\Delta H_S = -684.1$ [kJ/mol] $\Sigma KCl(s) \rightarrow (H_2O) \rightarrow K^+(aq) + Cl^-(aq)$...endothermic, heat take-up $\Delta H_S = +17.1$ [kJ/mol]
 - i) most non-ionic bonds are endothermic (lattice energy smaller than in ionic bonds);
 - i) most gases are exothermic (no lattice present to be broken up);

E. of **Physical Change:** Processes involved in physical changes;

- ΔH_{frez} E. of **Freezing:** The negative of the enthalpy of melting.
- ΔH_{vap} E. of **Vaporization:** The difference in enthalpy per mole between the vapor and liquid states of a substance; e.g.: $\Delta H_{\text{vap}} = H_{\text{vapor}} - H_{\text{liquid}}$ water: +40.7 [kJ/mol]
- ΔH_{mel} E. of **Melting:** The difference in enthalpy per mole between the solid and liquid states of a substance; e.g.: $\Delta H_{\text{mel}} = H_{\text{liquid}} - H_{\text{solid}}$ water: +6.01 [kJ/mol]
- ΔH_{sub} E. of **Sublimation:** The enthalpy change per mole of molecules when a solid changes into vapor; e.g.: $\Delta H_{\text{sub}} = H_{\text{vap}} - H_{\text{solid}}$ always endothermic (positive)

Electron E.: Uptake or release of energy by giving off, or accepting an electron.

- ΔH_{Ea} - **Electron Affinity:** The energy released when an electron is added to a gas-phase atom or ion of the elements forming an anion; the negative of the electron-gain enthalpy (see table below); e.g.:
 $\Delta H > 0$: $Ne(g) + e^- \rightarrow Ne^-(g)$ energy is needed to add an e $\Delta H_{\text{Ea}} = +29$ [kJ/mol] or [eV/atom]
 $\Delta H < 0$: $Fe(g) + e^- \rightarrow Fe^-(g)$ energy is released $\Delta H_{\text{Ea}} = -328$ [kJ/mol]
with the highest value in the upper right and the lowest values at the lower left of the periodic table.
- ΔH_{Ion} - **Ionization Energy:** The minimum energy required to remove an electron from the ground state of a gaseous atom, molecule, or ion. The second ionization energy is the ionization energy for removal of a second electron and has to be higher than the first; (ΔH always positive)
e.g.: 1st IE: $Na \rightarrow Na^+ + e^-$ $\Delta H_{\text{Ion}} = +496$ [kJ/mol] = [eV/atom]
2nd IE: $Ag^+ \rightarrow Ag^{2+} + e^-$ $\Delta H_{\text{Ion}} = +4563$ [kJ/mol]
I.E. increases from left to right across the period and decreases down a group in the periodic table.

Electron Gain E.: see electron enthalpy.

Lattice E.: The standard enthalpy change for the conversion of an ionic solid to a gas of ions.

ΔH_f° **Standard E. of Formation:** The heat change that results when 1 mole of compound is formed from its elements in their standard states, i.e.: 101.3[kPa] and 298.15[K] = 25[°C];

e.g.: $\Delta H^\circ = \Sigma n \Delta H_f^\circ(\text{products}) - \Sigma n \Delta H_f^\circ(\text{reactants})$

ΔH° **Standard E. of Reaction:** Enthalpy change of the reaction carried out under standard-state conditions.

ΔH_c° **Standard E. of Combustion:** The change in enthalpy per mole of substance when it burns (reacts with Oxygen) completely under standard-state conditions;

e.g.: C converts to CO₂, H to H₂O, N to N₂, S to SO₂;

- Entropy:** A direct measure of the randomness or disorder of a system.
- Q - Heat:** The amount of energy in form of heat;
 $Q = \Delta H \cdot n = C \cdot \Delta T$ [J]
C - Heat Capacity: The amount of heat required to raise the temperature of a given quantity of the substance by 1°C;
 i.e.: to raise 1g of water from 14.5°C to 15.5°C the energy of 4.184[J] (=1cal) is required: $C = m \cdot s$ [J/K]
H. of Dilution: The heat change associated with the hydration process - see chemistry liquid.
H. of Hydration: The heat change associated with the hydration process.
H. of Solution: see enthalpy of solution.
Specific H. Capacity: The heat capacity per gram.
 $s = Q/(m \cdot \Delta T)$ [N·m/(kg·K)] = [J/(kg·K)]
- Hess's Law:** When reactants are converted to products, the change in enthalpy is the same whether the reaction takes place in one step or in a series of steps.
- Molarity:** The number of moles of solute in one liter of solution.
- M. Heat of Fusion:** The energy [kJ] required to melt one mole of solid;
 e.g. 1[mol] at 101[kPa]: $H_2O(s) \rightarrow H_2O(l)$ $\Delta H_{Fus} = +6.01$ [kJ/mol]
- M. Heat of Condensation:** Equal to molar heat of vaporization but with opposite sign.
- M. Heat of Crystallization:** Amount of heat withdrawn from one mole of liquid at crystallization; equal to molar heat of fusion but with opposite sign.
- M. Heat of Sublimation:** The energy [kJ] required to sublime one mole of solid;
 $\Delta H_{Sub} = \Delta H_{Fus} + \Delta H_{Vap}$ [kJ/mol]
- M. Heat of Vaporization:** The energy [kJ] required to vaporize one mole of liquid; i.e.: the energy needed to liberate single molecules from the liquid pool (D=volumsarbeit); see Claudius-Clapeyron's equation;
 e.g. 1 mol of at 25°C at 1atm: $H_2O(l) \rightarrow H_2O(g)$ $\Delta H_{Vap} = +43.8$ [kJ/mol]
- the higher the intermolecular forces the lower the vapor pressure;
 - the lower the temperature the smaller ΔH_V ($T_{crit}: \Delta H_V = 0!$).
- Temperature:** 1) How hot or cold a sample is. 2) The intensive property that determines the direction in which heat will flow between two objects.
- Absolute T.:** Temperature scale with the lowest possible temperature: $T = 0K$ ($K = ^\circ C + 273.15$)
- Critical T.:** The temperature above which a gas will not liquefy.
- Thermochemistry:** The study of heat changes in chemical reactions.
- T. Equation:** An equation that shows both the species involved and enthalpy relations;
 e.g.: $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ $\Delta H = -890$ [kJ/mol]
- Thermodynamics:** The study of the transformation of energy from one form to another.
- 1st Law:** Energy can be converted from one form to the another, but cannot be created or destroyed.
- 2nd Law:** The entropy of the universe increases in a spontaneous process and remains unchanged in an equilibrium process.
- 3rd Law:** The entropy of a perfect crystalline substance is zero at the absolute zero of temperature.
- Van't Hoff Equation:** see stoichiometry - equilibrium.
- W -Work:** The energy expended during the act of moving an object against an opposing force; one joule of work is done when a force of 1 newton is exerted over a distance of 1 meter (compare power):
 $W = F \cdot d$ [N·m] = [J]
- Endergonic:** (Gk. endo, within; ergon, work) Describing a chemical reaction that requires energy to proceed; (compare exergonic).
- Exergonic:** (L. ex, out; ergon, work) Energy-yielding, as in a chemical reaction; applied to a "downhill" process; compare endergonic).

Molar Heat of Vaporization and Fusion of selected substances [kJ/mol]

Substance	Boiling Point [°C]	ΔH_{Vap} [kJ/mol]	Melting Point [°C]	ΔH_{Fus} [kJ/mol]
Argon (Ar)	-186	6.3	-190	1.3
Methane (CH ₄)	-164	9.2	-183	0.84
Ethyl ether (C ₂ H ₅ OC ₂ H ₅)	34.6	26.0	-116.2	6.90
Ethanol (C ₂ H ₅ OH)	78.3	39.3	-117.3	7.61
Benzene (C ₆ H ₆)	80.1	30.8	5.5	10,9
Water (H ₂ O)	25	43.8	0	6.01
Water (H ₂ O)	100	40.79		
Mercury (Hg)	357	59	-39	23.4

Standard bond energies of formation for common substances (25°C and 101.3kPa)

[kJ/mol]

Compound	ΔH_f°	Compound	ΔH_f°	Compound	ΔH_f°
C-C	+347	Br-Br	+190	N-N	+159
C=C	+619	Cl-Cl	+243	N=N	+418
C≡C	+812	F-F	+155	N≡N	+941
C-H	+413	H-Br	+364	O-Cl	+207
C-Cl	+326	H-Cl	+431	O-F	+184
C-F	+485	H-F	+565	O-H	+463
C-N	+293	H-H	+435	O-O	+143
C=N	+616	H-I	+297	O=O	+494
C≡N	+879	I-I	+149	P-Cl	+326
C-O	+335			P-H	+318
C=O	+707	N-Cl	+201	S-Cl	+276
C≡O	+1070			S-H	+364

Standard enthalpies of formation of common substances (25°C and 101.3kPa)

[kJ/mol]

Compound	ΔH_f°	Compound	ΔH_f°	Compound	ΔH_f°
AgCl(s)	-127.0	Cl(aq)	-167.4	I ₂ (g)	0
Al ₂ O ₃ (s)	-1669.8	Cl ₂ (g)	0	KCl(s)	-436.0
B ₃ H ₆ (s)	+73.2	CO(g)	-110.5	KClO ₃ (s)	-391.2
B ₂ O ₃ (s)	-1273.5	CO ₂ (g)	-393.5	KClO ₄ (s)	-435.1
BaCO ₃ (s)	-1218	COCl ₂ (g)	-223	LiAlH ₄ (s)	-100
BaO(s)	-588.1	CS ₂ (l)	+87.86	MgO(s)	-601.83
C (atomic)	+716.7	Cu ²⁺ (aq)	+64.4	MgCO ₃ (s)	-1112.9
C (diamond)	+1.9	CuO(g)	-155.2	NaCl(s)	-411.0
CaCO ₃ (s)	-1206.9	F ₂ (g)	0	NaOH(s)	-425.6
CaO(s)	-635.5	Fe ²⁺ (aq)	-87.9	N ₂ (g)	0
Ca(OH) ₂ (s)	-986.59	Fe ₂ O ₃ (s)	-822.2	NF ₃ (g)	-113
CF ₄ (g)	-913.4	H⁺(aq)	0.0	NH ₃ (g)	-46.19
CH ₄ (g)	-74.85	H ₂ (g)	0	NH ₃ (l)	-66.9
C ₂ H ₂ (g)	+226.7	HBr(g)	-36.2	NH ₄ NO ₃ (s)	-365.1
C ₂ H ₄ (g)	+52.3	HCl(g)	-92.3	NO(g)	+90.37
C ₂ H ₆ (g)	-84.68	HCN(g)	+130.5	NO ₂ (g)	+33.8
C ₃ H ₈ (g)	-103.8	HF(g)	-269	N ₂ O(g)	+81.6
C ₄ H ₁₀ (g)	-126.1	HgBr ₂ (s)	-169	N ₂ O ₄ (g)	+9.16
C ₅ H ₁₂ (g)	-146.4	HI(g)	+25.9	OH ⁻ (aq)	-229.95
C ₆ H ₆ (l)	+49.04	HNO ₃ (l)	-173.2	O ₂ (g)	0
C ₆ H ₁₂ O ₆ (s)	-1268	H ₂ O(g)	-241.8	O ₃ (g)	142.2
CH ₃ Cl(l)	-132	H ₂ O(l)	-285.9	PH ₃ (g)	+9.25
CH ₃ NH ₂ (g)	-28	H ₂ O ₂ (l)	-187.6	SO ₂ (g)	-296.9
CH ₃ OH(g)	-201.2	H ₂ S(g)	-20.2	SO ₃ (g)	-395.2
CH ₃ OH(l)	-238.6			Zn ²⁺ (aq)	-152.3
C ₂ H ₅ OH(l)	-277.6			ZNO(s)	-348.0

all other solid elements, like Ag(s), Al(s), Hg(l), etc. have values of $\Delta H_f^\circ = 0$.

Electron affinity of common elements (theoretical values)

[kJ/mol]

	I	II			III	IV	V	VI	VII	VIII
1	H -73									He (+21)
2	Li -60	Be (+240)			B -27	C -122	N 0	O -141	F -328	Ne (+29)
3	Na -53	Mg (+230)			Al -43	Si -134	P -72	S -200	Cl -349	Ar (+35)
4	K -48	Ca (+156)			Ga -29	Ge -116	As -77	Se -195	Br -325	Kr (+39)
5	Rb -47	Sr (+168)			In -29	Sn -121	Sb -101	Te -190	I -295	Xe (+41)
6	Cs -45	Ba (+52)			Ti -29	Pb -35	Bi -91	Po -183	At -270	Rn (+41)

Glossary - Chemistry (ElectroChemical Series)

Battery: An electrochemical cell, or often several electrochemical cells connected in series, that can be used as a source of direct electric current at a constant voltage.

Conductor: Any material through which charge easily flows (carries electric current) when subject an external electrical force; are also good heat conductors as well; electrons in the outer atomic shell are loose (valence shell).

C. Band: An incompletely filled band of orbitals in a solid.

Conductivity: The intrinsic property of a substance to conduct electric current; reciprocal to resistivity.

Metallic C.: An externally imposed potential difference forces the electrons in the conduction band, which are the carrier of the electric current, to move according to the gradient imposed.

Electrolytic C.: Carrier of the electric current are ions, results in a net transport of mass.

Corrosion: The deterioration of metals by an electrochemical process.

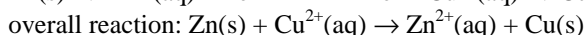
Daniell Cell: Consists of Cu- and Zn electrodes dipping into solutions of CuII-sulfate and Zn-sulfate; the 2 solutions make contact through a porous cup, which allows ions to pass through, to complete the circuit.

Electrochemistry: The branch of chemistry that deals with the use of chemical reactions to produce electricity, the relative strengths of oxidizing and reducing agents, and the use of electricity to produce chemical change.

Electrochemical Cell (Voltaic Cell): A system consisting of two electrodes in contact with an electrolyte.

Galvanic.C. An electrochemical cell used to produce electricity by means of a spontaneous redox reaction.

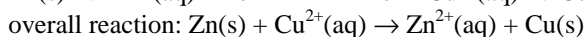
Here anode (-), cathode (+) to collect, give off the electrons of the spontaneously occurring reaction; e.g.: redox-reaction at the electrodes resulting in two half-cell reactions:



Electric C.: An electrochemical cell in which an electric current is used to cause chemical change;

Here anode (+), cathode (-) to start a reaction by imposing an external force during electrolysis.

e.g.: Zn electrode (anode) + Cu electrode (cathode) produces an electric current if emerged in an electrolyte; resulting in two half reactions:

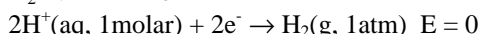
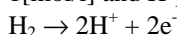


Electrode: A metallic conductor that makes contact with an electrolyte in an electrochemical cell - see there.

Anode: (Gk: an, up) The electrode at which oxidation occurs; attracts anions; e.g.: Cl⁻.

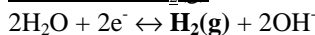
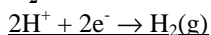
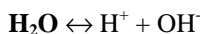
Cathode: (Gk: cat, down) The electrode at which reduction occurs; attracts cations; e.g.: Na⁺.

SHE - Standard Hydrogen E.: A H-electrode that is in its standard state (H⁺ ions at concentration 1 [mol/l] and H-pressure 101 [kPa]) and is defined as having E° = 0:



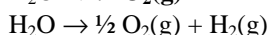
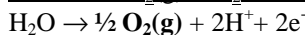
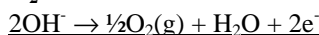
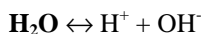
Electrolysis: 1) A process in which a chemical change is produced by passing an electric current through a liquid. 2) The process of driving a reaction in a non-spontaneous direction by passing an electric current through a solution: e.g.: electrolysis in aqueous solution of Na₂SO₄:

cathode-reaction:



Net result of these two half reactions:

anode reaction:



- electrolysis in aqueous solution of NaCl yields: $2\text{H}_2\text{O} + 2\text{Na}^+ + 2\text{Cl}^- \leftrightarrow \text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) + 2\text{Na}^+ + 2\text{OH}^-$
- electrolysis in aqueous solution of CuSO₄ yields; $\text{H}_2\text{O} + \text{Cu}^{2+} + \text{SO}_4^{2-} \leftrightarrow \text{Cu(s)} + \text{H}_2(\text{g}) + 2\text{H}^+ + \text{SO}_4^{2-}$

Electrolyte: 1) An ionically conducting medium. 2) A substance that, when dissolved in water, results in a solution that can conduct electricity; see table below.

E. Conduction: see conductor.

E. Rule: For a net potential of zero, the positive and negative charges must add up to zero; a solution must contain essentially as many anionic as cationic charges; see chemistry - liquid.

Non-E.: Is a solution in which no proportion of the solute molecules are ionized, hence does not conduct electricity; e.g.: $C_6H_{12}O_6(aq)$

Strong E.: Is a solution in which a large proportion of the solute molecules are ionized (complete dissociation into ions); e.g.: $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$ $1 \text{ mol } K_2SO_4 \rightarrow 3 \text{ mol ions}$

Weak E.: Is a solution in which only a small proportion of the solute molecules are ionized (partly dissociation into ions); e.g.: $CH_3COOH(aq)$ $1 \text{ mol } NaCl \rightarrow 2 \text{ mol ions}$

Electronegativity: The ability of an atom to attract electron toward itself in a chemical bond.

Faraday's Law of Electrolysis: The amount (in moles) of product formed by an electric current is chemically equivalent to the amount (in moles) of electrons supplied:

$$q = F_{EQ} \cdot z \cdot n = F_{EQ} \cdot z \cdot m/M \quad [C]$$

F. Constant: The charge per mole of electrons;

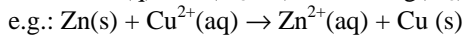
$$F = 6.022 \cdot 10^{23} [1/mol] \times 1.6022 \cdot 10^{-19} [C] = 96485 [C/mol]$$

Half Cell Reaction: Oxidation and reduction reactions at the electrodes - see electrochemical cell.

Redox Couple: Consists of the oxidized and reduced species taking part in the half reaction; e.g.: Red (Zn) / Ox (Zn^{2+})

Nernst Equation: The EQ expressing the cell potential in terms of the concentrations of the reagents taking part in the cell reaction;

$$E = E^\circ - \ln(q) \cdot R \cdot T / (N_{OX} \cdot F) = E^\circ - \log(K_C) \cdot 0.05916 / N_{OX} \quad [V]$$



$$= 1.1 - \log(c_{(Zn^{2+})} / c_{(Cu^{2+})}) \cdot 0.0592 / 2$$

Redoxpotential of any half reaction of a metal

$$E = E^\circ + \log(M^{n+}) \cdot 0.05916 / N_{OX} \quad [V]$$

V - Potential (voltage): The electric (pressure) potential energy per amount of charge, measured in volts, see physics electromagnetics.

$$V = PE/q \quad [J/(A \cdot s)] = [J/C] = [V]$$

E° - **Standard Cell P.:** The cell potential when the concentration of each type of ion taking part in the cell reaction is 1 [mol/l] and all the gases are at 1 [atm] pressure. E° is the sum of its two standard electrode potentials: $E^\circ = E^\circ(\text{cathode}) + E^\circ(\text{anode})$ [V]; see table below.

$E^\circ(\text{ox/red})$ - **Standard Electrode P.:** The contribution of an electrode to the standard cell potential see redox couple (half cell reaction).

Standard Reduction P.: The voltage measured as a redox reaction occurs at the electrode when all solutes are 1 [mol/l] and all gases are at 101 [kPa].

- Rules for SRP:
- i) all E° values of the half reactions are read from left to right
 - i) the more positive the greater the tendency to be reduced
 - i) half reactions are reversible; any electrode can act either as an anode or as a cathode
 - i) under STP conditions any species on the left of a given halfreaction will react spontaneously with a species that appears on the right above it (diagonal rule)
 - i) a change in stoichiometric coefficient does not effect the value of E°
 - i) E° changes sign whenever a halfreaction is reversed

Classification of solutes in aqueous solution

Strong e-lyte	Weak e-lyte	Non-e-lyte
HCL	CH ₃ COOH	(NH ₂) ₂ CO (urea)
HNO ₃	HF	CH ₃ OH (methanol)
HClO ₄	HNO ₂	C ₂ H ₅ OH (glucose)
H ₂ SO ₄	NH ₃	C ₆ H ₁₂ O ₆ (glucose)
NaOH	H ₂ O	C ₁₂ H ₂₂ O ₁₁ (sucrose)
Ba(OH) ₂		
ionic compounds		

Standard reduction potentials at 25[°C] ECS

Red	oxidizing agent (removes e ⁻ - reduces itself)	reducing agent (supplies e ⁻ - oxidizes itself)	Ox	E° [V]
weak	$F_2(g) + 2e^- \rightarrow 2F^-(aq)$		strong	+2.87
	$O_3(g) + 2H^+(aq) + 2e^- \rightarrow O_2(g) + H_2O(l)$			+2.07
	$Co^{3+}(aq) + e^- \rightarrow Co^{2+}(aq)$			+1.82
	$2H_2O_2(aq) + 2H^+(aq) + 2e^- \rightarrow 2H_2O$			+1.77
	$PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O(l)$			+1.70
	$Ce^{4+}(aq) + e^- \rightarrow Ce^{3+}(aq)$			+1.61
	$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O$			+1.51
	$Au^{3+}(aq) + 3e^- \rightarrow Au(s)$			+1.50
	$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$			+1.36
	$CrO_2^{2-}(s) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$			+1.33
	$MnO_2(s) + 4H^+(aq) + 2e^- \rightarrow Mn^{2+}(aq) + 2H_2O(l)$			+1.23
	$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$			+1.23
	$Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$			+1.07
	$NO_3^-(aq) + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(l)$			+0.96
	$2Hg^{2+}(aq) + 2e^- \rightarrow 2Hg_2^{2+}(aq)$			+0.92
	$Hg_2^{2+}(aq) + 2e^- \rightarrow 2Hg(l)$			+0.85
	$Ag^+(aq) + e^- \rightarrow Ag(s)$			+0.80
	$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$			+0.77
	$O_2(g) + 2H^+(aq) + 2e^- \rightarrow H_2O_2(aq)$			+0.68
	$MnO_4^-(aq) + 2H_2O + 3e^- \rightarrow MnO_2(s) + 4OH^-(aq)$			+0.59
	$I_2(s) + 2e^- \rightarrow 2I^-(aq)$			+0.53
	$O_2(g) + 2H_2O + 4e^- \rightarrow 4OH^-(aq)$			+0.40
	$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$			+0.34
	$AgCl(s) + e^- \rightarrow Ag(s) + Cl^-(aq)$			+0.22
	$SO_4^{2-}(aq) + 4H^+ + 2e^- \rightarrow SO_2(g) + 2H_2O(g)$			+0.20
	$Cu^{2+}(aq) + e^- \rightarrow Cu^+(aq)$			+0.15
	$Sn^{4+}(aq) + 2e^- \rightarrow Sn^{2+}(aq)$			+0.13
	$2H^+(aq) + 2e^- \rightarrow H_2(g)$			0.00
	$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$			-0.13
	$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$			-0.14
	$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$			-0.25
	$Co^{2+}(aq) + 2e^- \rightarrow Co(s)$			-0.28
	$PbSO_4(s) + 2e^- \rightarrow 2Pb(s) + SO_4^{2-}(aq)$			-0.31
	$Cd^{2+}(aq) + 2e^- \rightarrow Cd(s)$			-0.40
	$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$			-0.44
	$Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$			-0.74
	$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$			-0.76
	$2H_2O(l) + 2e^- \rightarrow 2H_2(g) + 2OH^-(aq)$			-0.83
	$Mn^{2+}(aq) + 2e^- \rightarrow Mn(s)$			-1.18
	$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$			-1.66
	$Be^{2+}(aq) + 2e^- \rightarrow Be(s)$			-1.85
	$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$			-2.37
	$Na^+(aq) + e^- \rightarrow Na(s)$			-2.71
	$Ca^{2+}(aq) + 2e^- \rightarrow Ca(s)$			-2.87
	$Sr^{2+}(aq) + 2e^- \rightarrow Sr(s)$			-2.89
	$Ba^{2+}(aq) + 2e^- \rightarrow Ba(s)$			-2.90
	$K^+(aq) + e^- \rightarrow K(s)$			-2.93
strong	$Li^+(aq) + e^- \rightarrow Li(s)$		weak	-3.05

Glossary - Chemistry (Solutions / RedOx)

Disproportionation Reaction: see Redox R.

Oxidation: 1) Combination with oxygen. 2) A reaction in which an atom, ion, or molecule loses an electron;
e.g.: $\text{Ca(s)} \rightarrow \text{Ca}^{2+}(\text{s}) + 2\text{e}^-$ (represents the Ca in CaO).

Oxidation corresponds to an increase of oxidation number.

O. Agent: A substance that can accept electrons from another substance or increase the oxidation number in another substance (being oxidized); where the substance itself is reduced;

e.g.: O_2 , O_3 , MnO_4^- , Fe^{3+}

N_{ox} - O. Number: The effective charge on an atom in a compound, calculated according to a set of rules. An increase in ON. corresponds to oxidation, and a decrease to reduction.

O. Reaction: The half-reaction that involves the loss of electrons;

O. Rules: Rules to assign ON; see also table below.

1. In free elements, each atom has an ON. of (0); e.g.: H_2 , He, Br_2 , Na, Be, K, O_2 , P_4 , S_8 , etc.
2. Single-atomic ions; ON. is equal to the charge on the ion; e.g. Na^+ (+1), Mg^{2+} (+2), Fe^{3+} (+3), etc.
3. In neutral molecules, the sum of the ON. of all the atoms must be 0; in polyatomic ions, the sum of ON. of all atoms in the ion must be equal to the net charge of the ion; e.g.: NH_4^+ (-3) + 4(+1) = (+1).
4. The most negative element **F** has an ON. of (-1); the 2nd-most negative, O usually (-2).
5. **H** has an ON. of (+1), e.g. HCl; when bonded to metals, the O.N. is (-1), e.g.: LiH, NaH, CaH_2 .
6. Compounds consisting of non-metallic atoms, O.N. is equal to the individual ionic net charge; e.g.: PCl_3 (+3) + 3(-1) = (0)

RedOx Reaction: A reaction in which there is either a transfer of electrons or a change in the oxidation numbers of the substances taking part in the reaction. Oxidation and reduction takes place simultaneously, because an electron that is lost by one atom is accepted by another. Oxidation-reduction reactions are important means of energy transfer in living systems. e.g.: $\text{Ca(s)} + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{CaO(s)}$;

Balancing RedOx R.: Balancing Redox-Reactions is done with the ion-electron method:

1. Write unbalanced EQ for the reaction in ionic form;
2. Separate the equation into two-half reactions;
3. Balance the atoms other than O and H in each half reaction separately;
4. For reactions in acidic, basic media, add H_2O to balance the O atoms and H^+ to balance H atoms;
5. Determine N_{OX} , add electrons to one side of each half-reaction to balance the charges; if necessary balance electrons by multiplying one or both half reactions by an appropriate coefficient;
6. Add the two half reactions together and balance the final EQ by inspection (e cancel on both sides);
7. Verify that the EQ contains the same types, masses, numbers of atoms, and charges on both sides;

Types of Redox R.:

Combination R.: Any combustion reaction that involves elemental oxygen or other e.g.: $\text{A} + \text{B} \rightarrow \text{C}$

Decomposition R.: A breakdown reaction of a compound into two or more components;

e.g.: $\text{C} \rightarrow \text{A} + \text{B}$

Displacement R.: An ion, or atom, in a compound is replaced by an ion, or atom of another one;

e.g.: $\text{A} + \text{BC} \rightarrow \text{AC} + \text{B}$

- **Hydrogen D.:** All group-I and some group-II elements (Ca, Sr, Ba) will displace H from water;
- **Metal D.:** Metals ranking higher in their period will displace the lower ranked elements.
- **Halogen D.:** The power of elements in group-7 decrease as we move down from F to I; molecular F can replace Cl, Br, and I etc.

Disproportion R.: A RedOx reaction in which a single element is simultaneously oxidized and reduced;
e.g.: $2\text{Cu}^+(\text{aq}) \rightarrow \text{Cu(s)} + \text{Cu}^{2+}(\text{aq})$.

Komproportion R.: The opposite reaction of disproportion; in which two compounds react to form a product; e.g.: $2\text{MnO}_4^-(\text{s}) + 3\text{Mn}^{2+}(\text{aq}) \rightarrow 5\text{MnO}_2(\text{aq})$

Reduction: (L. reductio, bringing back) 1) The removal of oxygen form (bringing back a metal from its oxide) or the addition of hydrogen to a compound. 2) A reaction in which an atom, an ion, or a molecule gains an electron; reduction takes place simultaneously with oxidation;

e.g.: $\frac{1}{2}\text{O}_2(\text{g}) + 2\text{e}^- \rightarrow \text{O}^{2-}$ (represents the O in CaO).

Reduction corresponds to a decrease in oxidation number.

R. Agent: A substance that can donate electrons to another substance or decrease the oxidation numbers in another substance (being reduced), where the substance itself is oxidized; e.g.: H_2 ; H_2S , SO_3^{2-}

R. Reaction: The half-reaction that involves the gain of electrons; e.g.: $\text{S(s)} + 3\text{Fe}_2(\text{g}) \rightarrow \text{SF}_6$

Oxidation Number of elements in their compounds:

	I	II										III	IV	V	VI	VII	VIII	
1	H +1 -1																He	
2	Li +1	Be +2										B +3	C +4 +2 -4	N +5 +4 +3 +1 -3	O +2 -1/2 -1 -2	F -1	Ne	
3	Na +1	Mg +2										Al +3	Si +4 +4	P +5 +3 -3	S +6 +4 +3 -2	Cl +7 +6 +5 +4 +3 +1 -1	Ar	
4	K +1	Ca +2	Sc +3	Ti +4 +3 +2	V +5 +4 +3 +2	Cr +6 +5 +4 +3 +2	Mn +7 +6 +4 +3 +2	Fe +3 +2	Co +3 +2	Ni +2	Cu +2 +1	Zn +2	Ga +3	Ge +4 -4	As +5 +3 -3	Se +6 +4 -2	Br +5 +3 +1 -1	Kr +4 +2
5	Rb +1	Sr +2	Y +3	Zr +4	Nb +5 +4	Mo +6 +4 +3	Tc +7 +6 +4	Ru +8 +6 +4 +3	Rh +4 +3 +2	Pd +4 +2	Ag +1	Cd +2	In +3	Sn +4 +2	Sb +5 +3 -3	Te +6 +4 -2	I +7 +5 +3 +1 -1	Xe +6 +4 +2
6	Cs +1	Ba +2	La +3	Hf +4	Ta +5	W +6 +4	Re +7 +6 +4	Os +8 +6 +4	Ir +4 +3	Pt +4 +2	Au +3 +1	Hg +2 +1	Tl +3 +1	Pb +4 +2	Bi +5 +3	Po +2	At -1	Rn
7	Fr +1	Ra +2																

- The sum of the oxidation numbers of all the atoms in the species is equal to its total charge.
- For the atoms in element form it is **0**.
- For elements of group

I = +1

II = +2

III = +3 (except B)

IV = +4 (except Si and C)

-1 in combination with metals

- For **H** +1 in combination with nonmetals
- For **F** is always **-1**.
- For **O** -2 unless combined with **F**

-1 in peroxides O_2^{2-}

-1/2 in superoxides O_2^-

-1/3 in ozonides O_3^-

Glossary - Chemistry (Solutions / Acid and Bases)

Acid-Base Concepts: Acids are molecules or ions that can donate protons to other molecules; in aqueous solutions generate cations by proton transfer (H^+) to water molecules (H_3O^+); bases are molecules or ions that can accept protons from other molecules; in aqueous solutions generate anions (OH^-) by accepting protons from water molecules; In terms of ionic charge, the acid in general is considered more positive than the base.

A.B. Indicator: see indicator.

A.B. reaction: A reaction between an acid and a base; reaction between conjugated acid-base pairs.

e.g.: $CH_3COOH(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + CH_3COO^-(aq)$

Arrhenius: Acid-base concept based on aqueous solutions only.

- **A. Acid:** A compound that contains hydrogen and releases hydrogen ions (H^+) in water; e.g.: HCL, CH_3COOH ; but not CH_4 .

- **A. Base:** A compound that produces hydroxide ions (OH^-) in water; e.g.: NaOH, NH_3 , but not Na.

Bronsted: Acids and bases description based on their proton acceptability (H^+) and donation (H^+).

- **B. Acid:** A substance capable of donating a proton; a source of H-ions, H^+ ;

e.g.: HCL, CH_3COOH ; HCO_3^- , NH_3

- **B. Base:** A substance capable of accepting a proton; a species to which hydrogen ions H^+ can bond; e.g.: OH^- , Cl^- , $CH_3CO_2^-$, HCO_3^- , NH_3 ;

Conjugate A.B. Pair: A bronsted acid and its conjugated base; i.e.: two substances related by gain or loss of an H^+ ion (proton); e.g.: HCL and Cl^- , NH_4^+ and NH_3 .

- **C. A.:** The bronsted acid formed when a bronsted base has accepted a proton;

e.g.: NH_4^+ is the conjugate acid of NH_3 .

- **C. B.:** The bronsted base formed when a bronsted acid has donated a proton;

e.g.: NH_3 is the conjugate acid of NH_4^+ .

Lewis A.B.: His theory is based upon the donation of electron pairs from one compound to another; hence does not necessarily exclude compounds which are not made partly of oxygen and hydrogen and is not limited to water as a solvent, instead any net ionic equation contains L-acid and L-bases.

- **L. Acid:** A substance that can accept a pair of electrons; e.g.: H^+ , Fe^{3+} , Cu^{2+} , BF_3 , etc.

- **L. Base:** A substance that can donate a pair of electrons; e.g.: OH^- , H_2O , NH_3 , CN^- , F^- , etc.

Polyprotic A.B.: A bronsted acid or base that can donate or accept more than one proton; see table.

- **Monoprotic A.:** A bronsted acid with one acidic hydrogen atom; each of the acid yields an H-atom; e.g.: $HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$.

- **Diprotic A.:** Each unit of the acid yields two H-atoms; e.g.: $K_{A1} > K_{A2}$

$H_2SO_4(aq) + H_2O(l) \rightarrow H_3O^+(aq) + HSO_4^-(aq)$

K_{A1} = completely dissociated

$HSO_4^-(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + SO_4^{2-}(aq)$

$K_{A2} = 1.3 \cdot 10^{-2}$

- **Triprotic A.:** These yield three H^+ ions; e.g.: $K_{A1} > K_{A2} > K_{A3}$

$H_3PO_4(aq) + H_2O(l) \rightarrow H_3O^+(aq) + H_2PO_4^-(aq)$

$K_{A2} = 7.6 \cdot 10^{-3}$

$H_2PO_4^-(aq) + H_2O(l) \rightarrow H_3O^+(aq) + HPO_4^{2-}(aq)$

$K_{A2} = 6.2 \cdot 10^{-8}$

$\Delta = 10^5!$

$HPO_4^{2-}(aq) + H_2O(l) \rightarrow H_3O^+(aq) + PO_4^{3-}(aq)$

$K_{A2} = 1.1 \cdot 10^{-12}$

$\Delta = 10^4!$

Weak A.B.: The stronger electronegativity, and the larger the element, the more likely an H-atom can be drawn away from it: 1) Acids and bases that are only partially ionized in aqueous solutions at normal concentrations. 2) Acids and bases for which K_a and K_b , respectively, are small compared to 1; e.g.: HF, CH_3COOH (weak acids); NH_3 , CH_3NH_2 (weak bases).

Acid: A substance that dissociates in water by yielding hydrogen ions (H^+), with subsequent hydration;

e.g.: $HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$.

K_A - **Ionization constant:** The equilibrium constant for the proton transfer of water(acid concentration);

e.g.: $HA + H_2O \leftrightarrow H_3O^+ + A^-$

$$K_A = K_C \cdot c_{(H_2O)} = c_{(H_3O^+)} \cdot c_{(A^-)} / c_{(HA)}$$

Binary A.: An acid that contains only two elements.

Carboxylic A.: An acid that contains the carboxyl group $-COOH$.

Strength of A.: $H-A =$ Hydrogen Acid bond.

- *Acid* - the more polar the $H-A$ bond (due to a larger difference of electronegativity), the stronger the acid of the same period; e.g.: $H_2O < HF < HCl < HBr < HI$.
- The weaker the $H-A$ bond (due to greater atomic radii), the stronger the acid; this effect is dominant for acids of the same group; e.g.: $HF < HCl < HBr < HI$
- *OxoAcid* - the greater the number of O-atoms attached to the central atom (the greater its oxidation number (N_{OX}), the stronger the acid; e.g.: $HClO < HClO_2 < HClO_3 < HClO_4$
- If the same number of O-atoms are attached to the central atom (= greater EN of the central atom), the stronger the acid; e.g.: $HClO < HBrO < HIO$

Amphoteric: The ability to react with both acids and bases; e.g.: H_2O ; HCO_3^- , NH_3 , HSO_4^- , see oxides.

Anylate: see titration.

Arrhenius: see acid-base.

Base: A substance that dissociates in water by yielding hydroxide ions (OH^-) with subsequent hydration;

e.g.: $NaOH(s) \rightarrow H_2O \rightarrow Na^+(aq) + OH^-(aq)$

K_B - **B. Ionisation Constant:** Equilibrium constant for the proton transfer from water(base ionization);

e.g.: $B + H_2O \leftrightarrow BH^+ + OH^-$

$$K_B = c_{(BH^+)} \cdot c_{(OH^-)} / c_{(B)}$$

Strength of A.B.: The stronger the base the weaker its conjugate acid and vice versa.

Buffer: A solution that resist any change in pH when small amounts of acid or base are added; i.e. a weak acid or base with its conjugated counterpart in relative high concentrations; see Henderson-Hasselbalch.

Acid B.: Stabilizes solutions at $pH < 7$; e.g.: a solution containing both CH_3CO_2H and $CH_3CO_2^-$.

Base B.: Stabilizes solutions at $pH > 7$; e.g.: a solution containing NH_3 and NH_4^+ .

B. Capacity: The amount of hydronium ions (H_3O^+) or hydroxide ions (OH^-) that can be added to a buffer solution without changing its pH by more than 1 unit.

B. Range: The pH range over which a buffer solution is most effective.

B. Solution: A solution of a weak acid or base and its salt; both components must be present. The solution has the ability to resist changes in pH upon the addition of small amounts of either acid or base.

α - **% Dissociation:** Overall concentration of a weak acid.

$$\alpha = 100 \cdot c_{(A^-)} / c_{(0)} = 100 \cdot c_{(A^-)} / (c_{(HA)} + c_{(A^-)}) \quad [-]$$

$$K_A = \alpha^2 \cdot c_{(0)} / (1 - \alpha) \quad [mol/l]$$

$c_{(0)}$, initial concentration [mol/l]

$c_{(A^-)}$, conc. conj. base [mol/l]

$c_{(HA)}$, concentr. of acid [mol/l]

Equivalence Point: The point at which the acid has completely reacted (neutralized) with the base - see titration.

Henderson-Hasselbalch Equation: An approximate equation for

estimating the pH of a solution containing a weak conjugate acid and base, and specifically that of a buffer solution.

$$pH = \log(K_A) + \log(c_{(A^-)}) - \log(c_{(HA)}) \quad [-]$$

$$pH = pK_A + \log(c_{(A^-)} / c_{(HA)}) \quad [-]$$

K_A , ioniz. const. [mol/l]

$c_{(A^-)}$, conc. of conj. base [mol/l]

$c_{(HA)}$, concentr. of acid [mol/l]

Hydroxyl: see ion.

Indicator: A weak organic acid or base that changes color when it shift from acid to its base form; i.e. acid-base neutralization (acid-base indicator) or from its oxidized to its reduced form (a redox indicator); e.g.:

$HInd(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + Ind^-(aq)$ where $c_{(H_3O^+)}$ is the H^+ concentration (pH) of the anylate.

$$K_{(Ind)} / c_{(H_3O^+)} = c_{(HInd)} / c_{(Ind^-)} \quad [mol/l]$$

if $= c_{(HInd)} / c_{(Ind^-)} > 1$, color of acid (Hind) predominates

if $= c_{(HInd)} / c_{(Ind^-)} < 1$, color of conjugate base (Ind^-) predominates

Preconditions are: both 1) HIn and In^- have to be water soluble, 2) HIn and In^- have to separate colors, 3)

Indicator concentration $c_{(HIn)}$ has to be low, 4) $c_{(H_3O^+)}$ of indicator should measure only pH of solution.

Ion: An atom / molecule that has lost / gained one or more electrons, and thus becomes positively or negatively charged; i.e.: Al^{3+} (mono-atomic ion), SO_4^{2-} (poly-atomic ion); see chemistry-atom and molecule.

Common I. Effect: The shift on equilibrium caused by the addition of a compound having an ion in common with the dissolved substances e.g.:

acid and electrolyte of the acetate ion - CH_3COOH and CH_3COONa ;

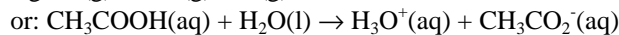
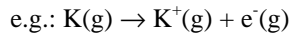
Hydration of I.: The attachment of water molecules to a central ion.

Hydronium I.: The ion H_3O^+ ; $H_5O_3^+$; $H_7O_5^+$;

Hydroxyl Group: An OH^- group; a negatively charged ion formed by the disassociation of a water molecule-compound that consists only of hydrogen and carbon atoms.

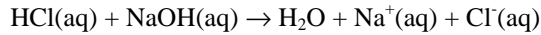
Oxoanion: An anion derived from an oxoacid; e.g.: HCO_3^- , CO_3^{2-} .

Ionization: 1) Conversion to cations by the removal of electrons. 2) The donation of a proton from a neutral acid molecule to a base with the formation of the conjugate base (an anion, in this instance) of the acid (dissociation into ions of a compound in solution);



Lewis: see acid-base.

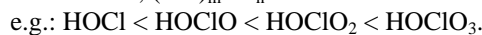
Neutralizing Reaction: A reaction between an acid and a base resulting in the formation of solute and a salt, where the acid provides the anions and the base the cations: e.g.:



Oxoacid: An acid that contains O, H and another central element; e.g.: H_2CO_3 , HNO_3 , HNO_2 , $HClO$;

- **Metallic OA.:** H-O-Z (Z stands for the metallic compound), showing lesser electronegativity than O, i.e.: H-O bond is more covalent, leading to hydroxide (OH⁻) separation when dissolved in water;
- **Non-Metallic-OA.:** H-O-Z (Z stands for the non-metallic compound), here the O-Z bond is more covalent, leading to hydrogen (H⁺) separation when dissolved in water; e.g.: HOI, HOBr, HOCl, ect.

Multiple OA.: Acidity increases with extra O-atoms in the compound (relocate electron-density), which facilitates H⁺ dissociation; $(HO)_mZO_n$



n = 0, weak acidity

n = 1, intermediate acidity

n = 2, strong acid

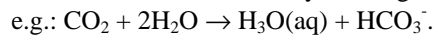
n = 3, very strong acid

Non-OA.: Acids in which the H-atom is not bounded to an O-atom; e.g.: HF, HCl, H₂S, etc. in which acidity increases moving down a group: $NH_3 < H_2O < HF$ or $PH_3 < H_2S < HCl$

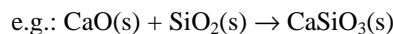
Oxide: see table below

Acidic O.: An oxide of nonmetallic elements (p-block) that reacts with water to give an acid;

- Soluble: React with water by forming acids; like CO_2 ; N_2O_5 , SO_2 , SO_3



- Insoluble: Acid oxides can react with basic oxides; $CaCO_3(s) \rightarrow T \rightarrow CaO(s) + CO_2(g)$

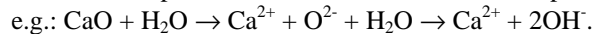


Amphoteric: The ability to react with both acids and bases; e.g.: H_2O ; HCO_3^- , NH_3 , HSO_4^- ,

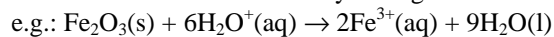
- **A. Oxide:** An oxide that exhibits both acidic and basic properties.

Basic O.: The oxides (Bronsted base) of metallic elements (s-block) are generally basic e.g.: MgO , Na_2O etc. except BeO .

- Soluble: Group-I and II elements; CaO , SrO , BaO possess ionic structure already in the solid phase;



- Insoluble: Become soluble only during a neutralization reaction;

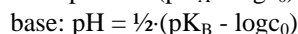
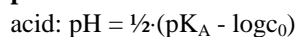


pH (potential hydrogen): The negative logarithm of the hydronium ion concentration in a solution, at 25°C;

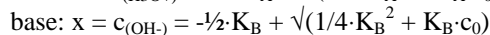


pH < 7 indicates an acidic solution, and pH > 7 a basic solution;

pH for weak acids/bases: simplification is applicable



pH for intermediate acids/base: no simplification applicable here:



pH-Curve: The graph of the pH of a reaction mixture

against volume of titrant added in an acid-base titration.

pOH (potential hydroxide) The negative logarithm of the

hydroxide ion concentration in a solution, at 25°C;

ionic product of pH and pOH reflects pK_W

$$pH = -\log(c_{(H_3O^+)})$$

$$K_C = c_{(H_3O^+)} \cdot c_{(OH^-)} / c_{(H_2O)}^2$$

$$K_W = 10^{-14} [\text{mol}^2/\text{l}^2] = 10^{-7} [\text{mol/l}]$$

see autoionization of water for K_W

K_A , acid ionization const. [-]

K_B , base ionization const. [-]

c_0 , initial concentration [mol/l]

$$pH = \frac{1}{2} \cdot pK_W + \frac{1}{2} \cdot (K_B + \log c_0)$$

$$pOH = -\log(c_{(OH^-)})$$

$$K_W = 10^{-14} [\text{mol}^2/\text{l}^2]$$

$$pH + pOH = 14$$

Salt: The product (other than water) of the reaction between an acid and a base; see chemistry liquid - solubility equilibria e.g.: $NaCl$, K_2SO_4 , etc.

- S. from strong bases and acids: pH of solution remains unaffected and oscillates around pH = 7;

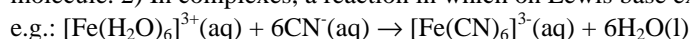
e.g. $NaCl$, KNO_3 , etc.

- S. from strong bases and weak acids: pH shifts to levels > 7; e.g.: KNO_2 , $NaCN$, $Ca(CH_3CO_2)_2$, etc.

- S. from weak bases and strong acids: pH shifts to levels < 7; e.g.: NH_4NO_3 , $FeBr_2$, $AlCl_3$, ect.

- S. from weak acids and bases: both acidic or basic solutions are possible, depending upon the acidic- or basic-character of the anions.

Substitution Reaction: 1) A reaction in which an atom (or a group of atoms) replaces an atom in the original molecule. 2) In complexes, a reaction in which one Lewis base expels another and takes its place;



Titration: The analysis of composition by measuring the volume of one solution (titrant) needed to react with a given volume of another solution (the analyte).

T. involving a strong acid and base: Equivalence point oscillates around pH 7 (steep rise at EP).

T. i.a. weak acid and strong base: Equivalence point oscillates around levels pH 7 - 14.

T. i.a. weak base and strong acid: Equivalence point oscillates around levels pH 1 - 7.

T. i.a. a weak acid and base: Equivalence point oscillates around pH 7, with a flat EP, hard to determine.

T. **Curve:** The sigmoidal profile of pH (analyte) over the volume of added titrant or level of titrant τ .

Analyte: The solution of unknown concentration in a titration.

Equivalence Point (stoichiometric point): The stage in a titration

when exactly the right volume of solution needed to complete

the reaction has been added; $n(\text{An})_{\text{An}} = n(\text{Ti})_{\text{Ti}}$

$$c_{\text{An}} \cdot V_{\text{An}} = c_{\text{Ti}} \cdot V_{\text{Ti}}$$

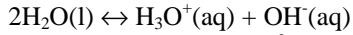
Titration: The solution of known concentration added from a buret in a titration

n , molar amount [mol]

V , volume of probes [l]

c , molar concentration [mol/l]

K_W - **Water Autionization Constant:** The equilibrium constant for the auto-ionization of water,



$$K_W = c_{(\text{H}_3\text{O}^+)} \cdot c_{(\text{OH}^-)} = K_C \cdot c_{(\text{H}_2\text{O})}^2 = 10^{-14} [\text{mol}^2/\text{l}^2]$$

$$K_W = K_A \cdot K_B$$

$$pK_W = pK_A + pK_B$$

Representative elements and their highest oxidation states

I		II												III	IV	V	VI	VII	VIII
Li ₂ O	BeO													B ₂ O ₃	CO ₂	N ₂ O ₅		OF ₂	
Na ₂ O	MgO													Al ₂ O ₃	SiO ₂	P ₄ O ₁₀	SO ₃	Cl ₂ O ₇	
K ₂ O	CaO													Ga ₂ O ₃	GeO ₂	As ₂ O ₅	SeO ₃	Br ₂ O ₇	
Rb ₂ O	SrO													In ₂ O ₃	SnO ₂	Sb ₂ O ₅	TeO ₃	I ₂ O ₇	
Cs ₂ O	BaO													Ti ₂ O ₃	PbO ₂	Bi ₂ O ₅	PoO ₃	At ₂ O ₇	

s-block p-block

basic oxide
 acidic oxide
 amphoteric oxide

Common acids and bases in aqueous solution:

	acid/base	conjugate a/b	K_A ** K_B	pK_A * pK_B
a	PerChloric Acid	$\text{HClO}_4 + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{ClO}_4^-$		≈ -10
a	Hydroiodic Acid	$\text{HI} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{I}^-$		≈ -10
a	Hydrobromic Acid	$\text{HBr} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{Br}^-$		≈ -9
a	Hydrochloric Acid	$\text{HCl} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{Cl}^-$		≈ -6
a	Sulfuric Acid	$\text{H}_2\text{SO}_4 + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{HSO}_4^-$		≈ -3
	Hydrogen sulfate ion	$\text{HSO}_4^- + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{SO}_4^{2-}$	$1.3 \cdot 10^{-2}$	1.89
a	Nitric Acid	$\text{HNO}_3 + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{NO}_3^-$		-1.3
a	Hydronium	$\text{H}_3\text{O}^+ + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{H}_2\text{O}$	$1.0 \cdot 10^0$	-1.7
b	Urea	$\text{NH}_2\text{CONH}_2 + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{NH}_2\text{CONH}_3^+$	$1.3 \cdot 10^{-14}$	13.90
a	Trichloroacetic Acid	$\text{CCl}_3\text{CO}_2\text{H} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{CCl}_3\text{CO}_2^-$	$3.0 \cdot 10^{-1}$	0.52
a	Benzenesulfonic Acid	$\text{C}_6\text{H}_5\text{SO}_3\text{H} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{C}_6\text{H}_5\text{SO}_3^-$	$2.0 \cdot 10^{-1}$	0.70
a	Iodic Acid	$\text{HIO}_3 + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{IO}_3^-$	$1.7 \cdot 10^{-1}$	0.77
a	Sulfurous Acid	$\text{H}_2\text{SO}_3 + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{HSO}_3^-$	$1.6 \cdot 10^{-2}$	1.81
		$\text{HSO}_3^- + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{SO}_3^{2-}$	$5.6 \cdot 10^{-8}$	7.25
a	Phosphorous Acid	$\text{H}_3\text{PO}_3 + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{H}_2\text{PO}_3^-$	$1.6 \cdot 10^{-2}$	1.79
		$\text{H}_2\text{PO}_3^- + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{HPO}_3^{2-}$	$7.0 \cdot 10^{-7}$	6.15
a	Oxalix Acid	$\text{H}_2\text{C}_2\text{O}_4 + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{HC}_2\text{O}_4^-$	$5.9 \cdot 10^{-2}$	1.22
		$\text{HC}_2\text{O}_4^- + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{C}_2\text{O}_4^{2-}$	$5.4 \cdot 10^{-5}$	4.27
a	Chlorous Acid	$\text{HClO}_2 + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{ClO}_2^-$	$1.0 \cdot 10^{-2}$	2.00
a	Dichloric Acid	$\text{Cl}_2\text{HC}_2\text{O}_2\text{H} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{Cl}_2\text{HC}_2\text{O}_2^-$	$3.3 \cdot 10^{-2}$	1.48

a	Phosphoric Acid	$\text{H}_3\text{PO}_4 + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{H}_2\text{PO}_4^-$	$7.6 \cdot 10^{-3}$	2.12	
	Dihydrogen phosphate ion	$\text{H}_2\text{PO}_4^- + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{HPO}_4^{2-}$	$6.2 \cdot 10^{-8}$	7.21	
	Hydrogen Phosphate ion	$\text{HPO}_4^{2-} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{PO}_4^{3-}$	$1.0 \cdot 10^{-12}$	12.0	
a	Chloroacetic Acid	$\text{CH}_2\text{ClCO}_2\text{H} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{CH}_2\text{ClCO}_2^-$	$1.4 \cdot 10^{-3}$	2.85	
a	Arsenic Acid	$\text{H}_3\text{AsO}_4 + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{H}_2\text{AsO}_4^-$	$2.5 \cdot 10^{-4}$	3.60	
		$\text{H}_2\text{AsO}_4^- + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{HASO}_4^{2-}$	$5.6 \cdot 10^{-8}$	7.25	
		$\text{HASO}_4^{2-} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{AsO}_4^{3-}$	$3 \cdot 10^{-13}$	12.5	
a	Lactic Acid	$\text{CH}_3\text{CH}(\text{OH})\text{CO}_2\text{H} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{CH}_3\text{CH}(\text{OH})\text{CO}_2^-$	$8.4 \cdot 10^{-4}$	3.08	
a	Nitrous Acid	$\text{HNO}_2 + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{NO}_2^-$	$4.3 \cdot 10^{-4}$	3.37	
a	Hydrofluoric Acid	$\text{HF} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{F}^-$	$3.5 \cdot 10^{-4}$	3.45	
a	Formic Acid	$\text{HCO}_2\text{H} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{HCO}_2^-$	$1.8 \cdot 10^{-4}$	3.75	
a	Cyanic Acid	$\text{HOCN} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{OCN}^-$	$1.2 \cdot 10^{-4}$	3.9	
a	Benzoic Acid	$\text{C}_6\text{H}_5\text{CO}_2\text{H} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{C}_6\text{H}_5\text{CO}_2^-$	$6.0 \cdot 10^{-5}$	4.19	
b	Aniline	$\text{C}_6\text{H}_5\text{NH}_2 + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{C}_6\text{H}_5\text{NH}_3^+$	$4.3 \cdot 10^{-10}$		9.37
a	Stickstoffwasserstoff Acid	$\text{HN}_3 + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{N}_3^-$	$1.9 \cdot 10^{-5}$	4.7	
a	Acetic Acid (HAc)	$\text{CH}_3\text{CO}_2\text{H} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{CH}_3\text{CO}_2^-$	$1.8 \cdot 10^{-5}$	4.74	
b	Nicotinic Acid	$\text{C}_5\text{H}_4\text{NCO}_2\text{H} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{C}_5\text{H}_4\text{NCO}_2^-$	$1.4 \cdot 10^{-5}$	4.85	
b	Pyridine	$\text{C}_5\text{H}_5\text{N} + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{C}_5\text{H}_6\text{O}^+$	$1.8 \cdot 10^{-9}$		8.75
b	Hydroxalamine	$\text{NH}_2\text{OH} + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{NH}_2\text{OH}_2^+$	$1.1 \cdot 10^{-8}$		7.97
a	Propane Acid	$\text{C}_2\text{H}_5\text{CO}_2\text{H} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{C}_2\text{H}_5\text{CO}_2^-$	$1.3 \cdot 10^{-6}$		
a	Carbonic Acid	$\text{H}_2\text{CO}_3 + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{HCO}_3^-$	$4.3 \cdot 10^{-7}$	6.37	
	Hydrogen Carbonate ion	$\text{HCO}_3^- + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{CO}_3^{2-}$	$4.8 \cdot 10^{-11}$	10.3	
a	Hydrosulfuric Acid	$\text{H}_2\text{S} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{HS}^-$	$1.1 \cdot 10^{-7}$	7.0	
	Hydrogen Sulfide ion	$\text{HS}^- + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{S}^{2-}$	$1.0 \cdot 10^{-14}$	14.0	
	Water	$\text{H}_2\text{O} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{OH}^-$	$1.0 \cdot 10^{-14}$	7.0	7.0
a	Hypochlorous Acid	$\text{HClO} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{OCl}^-$	$3.0 \cdot 10^{-8}$	7.53	
b	Hydrazine	$\text{N}_2\text{H}_4 + \text{H}_2\text{O} \leftrightarrow \text{N}_2\text{H}_3^+ + \text{OH}^-$	$9.8 \cdot 10^{-7}$		6.0
b	Nicotine	$\text{C}_{10}\text{H}_{14}\text{N}_2 + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{C}_{10}\text{H}_{13}\text{N}_2^+$	$1.0 \cdot 10^{-6}$		5.98
b	Morphine	$\text{C}_{17}\text{H}_{19}\text{O}_3\text{N} + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{C}_{17}\text{H}_{20}\text{O}_3\text{N}^+$	$1.6 \cdot 10^{-6}$		5.79
b	Hydrazine	$\text{NH}_2\text{NH}_2 + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{NH}_2\text{NH}_3^+$	$1.7 \cdot 10^{-6}$		5.77
a	Hypobromous Acid	$\text{HBrO} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{OBr}^-$	$2.0 \cdot 10^{-9}$	8.69	
b	Pyridine	$\text{C}_6\text{H}_5\text{N} + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{C}_6\text{H}_5\text{NH}^+$	$1.5 \cdot 10^{-9}$		5.2
a	Bromous Acid	$\text{H}_3\text{BO}_3 + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{H}_2\text{BO}_3^-$		9.1	
a	Boric Acid	$\text{B}(\text{OH})_3 + 2\text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{B}(\text{OH})_4^-$	$7.2 \cdot 10^{-10}$	9.14	
a	Hydrocyanic Acid	$\text{HCN} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{CN}^-$	$4.9 \cdot 10^{-10}$	9.31	
b	Ammonia	$\text{NH}_3 + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{NH}_4^+$	$1.8 \cdot 10^{-5}$		4.74
a	Phenol	$\text{C}_6\text{H}_5\text{OH} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{C}_6\text{H}_5\text{O}^-$	$1.3 \cdot 10^{-10}$	9.89	
a	Hypoiodous Acid	$\text{HIO} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{IO}^-$	$2.3 \cdot 10^{-11}$	10.64	
b	Trimethylamine	$(\text{CH}_3)_3\text{N} + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + (\text{CH}_3)_3\text{NH}^+$	$6.5 \cdot 10^{-5}$		4.19
b	Methylamine	$(\text{CH}_3)_2\text{NH} + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + (\text{CH}_3)_2\text{NH}_2^+$	$3.6 \cdot 10^{-4}$		3.44
b	Dimethylamine	$(\text{CH}_3)_2\text{NH} + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + (\text{CH}_3)_2\text{NH}_2^+$	$5.4 \cdot 10^{-4}$		3.27
b	Ethylamine	$\text{C}_2\text{H}_5\text{NH}_2 + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{C}_2\text{H}_5\text{NH}_3^+$	$6.5 \cdot 10^{-4}$		3.19
b	Triethylamine	$(\text{C}_2\text{H}_5)_3\text{N} + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + (\text{C}_2\text{H}_5)_3\text{NH}^+$	$1.0 \cdot 10^{-5}$		2.99
a		$\text{H}_2\text{O}_2 + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{HO}_2^-$		11.6	
b		$\text{H}_2\text{BO}_3^- + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{H}^+$			1.3
b		$\text{HBO}_3^{2-} + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{H}^+$			0.2
b	Hydroxide ion	$\text{OH}^- + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{H}_2\text{O}$	$1.0 \cdot 10^0$		0.0
b	Hydride ion	$\text{H}^- + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{H}_2$			-1.7
b	Amide ion	$\text{NH}_2^- + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{NH}_3$			≈ -9
b		$\text{O}^{2-} + \text{H}_2\text{O} \leftrightarrow 2\text{OH}^-$			≈ -10

**) $K_A \cdot K_B = K_W = 1.0 \cdot 10^{-14}$ 1.3 · 10⁻¹⁴*) $pK_A + pK_B = pK_W = 14$