Chemistry

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http://denisemeeks.com/science/notebooks/notebook_chemistry.pdf

Chemistry: Plasma, Gas, Liquid, and Solid

state of matter	plasma	gas	liquid	solid
volume/ shape	variable	assumes container shape	definite volume; assumes container shape	definite volume and shape
density	low	low	high	high
compressibility	varies	very compress-ible	slightly compressible	virtually incompress-ible
molecular motion	very free motion	very free motion	freely slide past each other	vibrate about a fixed position

Chemistry: Matter (1)

Avogadro's number N_A = 6.022×10^{23}/mol

atomic mass unit = 1/12 the mass of an atom of carbon-12 = approximately 1.66×10^{-27} kg

elementary charge e = 1.6022×10^{-19} C

electron mass m_e = 9.109×10^{-31} kilograms

proton mass m_p = 1.673×10^{-27} kilograms

neutron mass m_n = 1.675×10^{-27} kilograms

Boltzmann's constant k = 1.381×10^{-23} joules/K

gas constant R = 8.314 joules/mol K

Chemistry: Matter (2)

atom: basic unit of a chemical element

isotope: each of two or more forms of the same element that contain equal numbers of protons but different numbers of neutrons in their nuclei

proton: stable subatomic particle occurring in all atomic nuclei, with positive electric charge to that of an electron, but of opposite sign

neutron: subatomic chargeless particle about the same mass as a proton, present in all atomic nuclei except ordinary hydrogen

electron: stable subatomic particle with a charge of negative electricity, acts as the primary carrier of electricity in solids

element: distinguished by its atomic number, the number of protons in the nuclei of its atoms

Chemistry: Matter (3)

molecule: group of atoms bonded together, representing the smallest fundamental unit of a chemical compound that can take part in a chemical reaction

compound: substance consisting of atoms or ions of two or more different elements in definite proportions joined by chemical bonds

isomer: different arrangements of the same atoms

atomic number: number of protons in an atomic nucleus

atomic weight: average mass of a chemical element, expressed in atomic mass units; the atomic weight of an element having more than one principal isotope is calculated both from the atomic masses of the isotopes and from the relative abundance of each isotope in nature

law of conservation of mass: matter can't be created or destroyed

Chemistry: Matter (4)

anion: negatively charged ion

cation: positively charged ion

complex ion: an ion containing a central metal cation bonded to one or more molecules or ions

physical property: can be measured and observed without changing the identity of a substance

chemical property: requires a chemical changes

law of definite proportions: different samples of the same compound always contain its constituent elements in the same proportion by mass

law of multiple proportions: masses of one element that combine with another are in ratios of small whole numbers

Chemistry: Matter (5)

<u>atomic mass unit (amu)</u>: a mass exactly equal to 1/12th the mass of on carbon-12 atom
<u>average atomic mass</u>: sum of the masses of its isotopes, each multiplied by its natural abundance (the decimal associated with percent of atoms of that element that are of a given isotope)
<u>mole</u>: the amount of a substance that contains as many elementary entities (atoms, molecules, or other particles) as there are atoms in 12 grams of the carbon-12 isotope
<u>atomic radius</u>: ½ the distance between two nuclei of adjacent metal atoms
<u>ionic radius</u>: the radius of a cation or anion
ionization energy: the minimum energy in kJ/mol required to remove an electron from a gaseous atom in its ground state

Chemistry: Matter (6)

<u>diagonal relationships</u>: similarities between pairs of elements in different groups and periods of the periodic table
<u>lattice energy</u>: the energy required to completely separate one mole of a solid ionic compound into gaseous ions
<u>electronegativity</u>: ability of an atom to attract toward itself the electrons in a chemical bond

Chemistry: Periodic Table (1)

hydrogen 1 H 1.0079																	helium 2 He 4.0026
lithium 3 Li 6.941	beryllium 4 Be 9.0122											boron 5 B 10.811	carbon 6 C 12.011	nitrogen 7 N 14.007	oxygen 8 O 15.999	fluorine 9 F 18.998	neon 10 Ne 20.180
sodium 11 Na 22.990	magnesium 12 Mg 24.305											aluminium 13 Al 26.982	silicon 14 Si 28.086	phosphorus 15 P 30.974	sulfur 16 S 32.065	chlorine 17 Cl 35.453	argon 18 Ar 39.948
potassium 19 K 39.098	calcium 20 Ca 40.078		scandium 21 Sc 44.956	titanium 22 Ti 47.867	vanadium 23 V 50.942	chromium 24 Cr 51.996	manganese 25 Mn 54.938	iron 26 Fe 55.845	cobalt 27 Co 58.933	nickel 28 Ni 58.693	copper 29 Cu 63.546	zinc 30 Zn 65.39	gallium 31 Ga 69.723	germanium 32 Ge 72.61	arsenic 33 As 74.922	selenium 34 Se 78.96	bromine 35 Br 79.904	krypton 36 Kr 83.80
rubidium 37 Rb 85.468	strontium 38 Sr 87.62		yttrium 39 Y 88.906	zirconium 40 Zr 91.224	niobium 41 Nb 92.906	molybdenum 42 Mo 95.94	technetium 43 Tc [98]	ruthenium 44 Ru 101.07	rhodium 45 Rh 102.91	palladium 46 Pd 106.42	silver 47 Ag 107.87	cadmium 48 Cd 112.41	indium 49 In 114.82	tin 50 Sn 118.71	antimony 51 Sb 121.76	tellurium 52 Te 127.60	iodine 53 I 126.90	xenon 54 Xe 131.29
caesium 55 Cs 132.91	barium 56 Ba 137.33	57-70 *	lutetium 71 Lu 174.97	hafnium 72 Hf 178.49	tantalum 73 Ta 180.95	tungsten 74 W 183.84	rhenium 75 Re 186.21	osmium 76 Os 190.23	iridium 77 Ir 192.22	platinum 78 Pt 195.08	gold 79 Au 196.97	mercury 80 Hg 200.59	thallium 81 Tl 204.38	lead 82 Pb 207.2	bismuth 83 Bi 208.98	polonium 84 Po [209]	astatine 85 At [210]	radon 86 Rn [222]
francium 87 Fr [223]	radium 88 Ra [226]	89-102 **	lawrencium 103 Lr [262]	rutherfordium 104 Rf [261]	dubnium 105 Db [262]	seaborgium 106 Sg [266]	bohrium 107 Bh [264]	hassium 108 Hs [269]	meitnerium 109 Mt [268]	ununnilium 110 Uun [271]	unununium 111 Uuu [272]	ununbium 112 Uub [277]	nihonium 113 Nh [284]	flerovium 114 Fl [289]	moscovium 115 Mc [288]	livermorium 116 Lv [293]	tennessine 117 Ts [294]	oganesson 118 Og [294]

*Lanthanide series	lanthanum 57 La 138.91	cerium 58 Ce 140.12	praseodymium 59 Pr 140.91	neodymium 60 Nd 144.24	promethium 61 Pm [145]	samarium 62 Sm 150.36	europium 63 Eu 151.96	gadolinium 64 Gd 157.25	terbium 65 Tb 158.93	dysprosium 66 Dy 162.50	holmium 67 Ho 164.93	erbium 68 Er 167.26	thulium 69 Tm 168.93	ytterbium 70 Yb 173.04
**Actinide series	actinium 89 Ac [227]	thorium 90 Th 232.04	protactinium 91 Pa 231.04	uranium 92 U 238.03	neptunium 93 Np [237]	plutonium 94 Pu [244]	americium 95 Am [243]	curium 96 Cm [247]	berkelium 97 Bk [247]	californium 98 Cf [251]	einsteinium 99 Es [252]	fermium 100 Fm [257]	mendelevium 101 Md [258]	nobelium 102 No [259]

(Image source: http://www.wpclipart.com/science/atoms_molecules/periodic_tables/periodic_table_of_elements_BW.png.html, public domain)

Chemistry: Periodic Table (2)

<u>Group 1, 1A</u>: Alkali Metals, Lithium Group, H, Li, Na, K, Rb, Cs, Fr, most electropositive, soft, can be sliced with a knife, oxidation, low ionization energies, number in their compounds is +1, low melting points, low densities, very reactive, never found in the pure state in nature
<u>Group 2, 2A</u>: Alkali Earth Metals, Beryllium Group, Be, Mg, Ca, Sr, Ba, Ra, oxidation number in the combined form is almost always +2, react with water
<u>Group 3, 3B</u>: Scandium Group, Sc, Y, Lu, Lr,
<u>Group 4, 4B</u>: Titanium Group, Ti, Zr, Hf, Rf
<u>Group 5, 5B</u>: Vanadium Group, V, Nb, Ta, Db
<u>Group 6, 6B</u>: Chromium Group, Cr, Mo, W, Sg

Chemistry: Periodic Table (3)

<u>Group 7, 7B</u>: Manganese Group, Mn, Tc, Re, Bh
<u>Group 8, 8B</u>: Iron Group, Fe, Ru, Os, Hs
<u>Group 9, 8B</u>: Cobalt Group, Co, Rh, Ir, Mt
<u>Group 10, 8B</u>: Nickel Group, Ni, Pd, Pt, Ds
<u>Group 11, 1B</u>: Coinage Metals, Copper Group, Cu, Ag, Au, Rg, less reactive than group 1A, usually found in uncombined natural state
<u>Group 12, 2B</u>: Volatile Metals, Zinc Group, Zn, Cd, Hg, Cn
<u>Group 13, 3A</u>: Icosagens, Boron Group, B, Al, Ga, In, Tl, Nh, boron is a metalloid, the rest are metals, these elements form many molecular compounds
<u>Group 14, 4A</u>: Crystallogens, Carbon group, C, Si, Ge, Sn, Pb, Fl, form compounds in both the +2 and +4 oxidation states

Chemistry: Periodic Table (4)

Group 15, 5A: Pnictogens, Nitrogen Group, N, P, As, Sb, Bi, Mc, nitrogen forms many compounds
Group 16, 6A: Chalcogens, Oxygen Group, O, S, Se, Te, Po, Lv, these elements form a large number of compounds with nonmetals
Group 17, 7A: Halogens, Fluorine Group, F, Cl, Br, I, At, Ts, nonmetals, very reactive, high ionization energies, positive electron affinities, form many molecular compounds among themselves
Group 18, 8A: Noble gases, Helium or Neon Group, He, Ne, Ar, Kr, Xe, Rn, Og, exist as monatomic species, very stable, no tendency to accept extra electrons, not involved in natural biological processes

Chemistry: Liquid and Gas Elements at Atmospheric Conditions

gas		liquid
2 He helium	1 H hydrogen H_2	35 Br bromine
8 Ne neon	7 nitrogen N_2	80 Hg mercury
18 Ar argon	8 oxygen O_2	87 Fr francium
36 Kr krypton	9 fluorine F_2	
54 Xe xenon	17 chlorine Cl_2	
86 Rn radon		

Chemistry: Properties of Non-Metals (1)

hydrogen: simplest element, atomic form consists only at very high temperatures, diatomic molecule, colorless, odorless, nonpoisonous gas, at 1 atm liquid hydrogen has a boiling point of 20.3 K, three important isotopes are atomic hydogen, deuterium, and tritium
carbon: found free in the form of diamond and graphite, is a component of natural gas, petroleum, and coal, combines with oxygen to form carbon dioxide and carbonate in limestone and chalk, forms carbides, cyanides, and oxides
nitrogen: 78% of air by volume, essential element of life, a component of proteins and nucleic acids, ammonia and hydrazine are important compounds

Chemistry : Properties of Non-Metals (2)

phosphorus: occurs most commonly in nature as phosphate rocks which are mostly calcium phosphate, forms hydrides, halides, and oxides
oxygen: most abundant element in Earth's crust, 21% of air by volume, diatomic molecule, building block of almost all biomolecules, forms oxides, peroxides, superoxides, and ozone
sulfur: occurs commonly in nature in elemental form, largest known reserves are found in sedimentary deposits, occurs in sulfide minerals, forms oxides and acids
halogens: toxic, reactive nonmetals, strong oxidizing agents, form halides and oxides

Chemistry: Properties of Transition Metals

iron: after aluminum, is the most abundant metal in Earth's crust, found in many ores
copper: rare element, found in nature in the uncombined state as well as in many ores

Chemistry: Chemical Reactions and Nuclear Reactions

chemical reactions	nuclear reactions
1. Atoms rearranged by breaking and forming of chemical bonds.	**1. Elements or isotopes of same element converted from one to another.**
2. Only electrons in atomic or molecular orbitals involved in the breaking and forming of bonds.	**2. Protons, neutrons, electrons, and other elementary particles may be involved.**
3. Reactions accompanied by absorption or release of relatively small amounts of energy.	**3. Reactions accompanied by absorption or release of tremendous amounts of energy.**
4. Rates of reaction influenced by temperature, pressure, concentrations, and catalysts.	**4. Rates of reaction normally are not affected by temperature, pressure, and catalysts.**

(Source: p. 910)

Chemistry: Nuclear Stability

- nuclei that contain 2, 8, 20, 50, 82, or 126 protons or neutrons are generally more stable than other nuclei
- nuclei with even numbers of both protons and neutrons are generally more stable
- all isotopes of the elements with atomic numbers higher than 83 are radioactive; all isotopes of technetium and promethium are also radioactive

protons	neutrons	number of stable isotopes
odd	**odd**	**4**
odd	**even**	**50**
even	**odd**	**53**
even	**even**	**157**

(Source: p. 913)

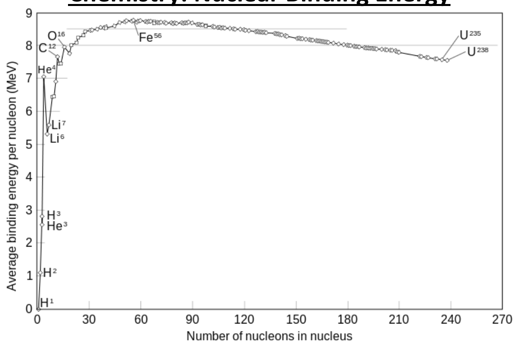
Chemistry: Bohr Hydrogen Atom and Electron Binding Energy

1. Electron orbits are discrete and non-radiating, and an electron may not remain between these orbits.
2. The energy change experienced by an electron changing from one orbit to another is quantized.
3. Classical mechanics does not hold when the electron is between orbits.
4. Angular momentum is quantized.

Bohr's equation $\bar{f} = R_\infty\left[\frac{1}{n_2^2} - \frac{1}{n_1^2}\right]$

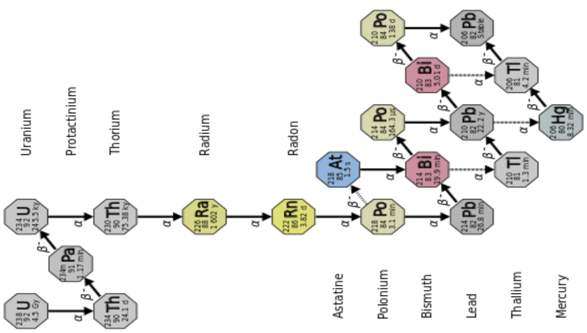
electron binding energy $E_{binding} = \frac{1}{2}mv^2 - \left|\frac{ke^2}{r}\right|$

Chemistry: Nuclear Binding Energy



(Image source: https://en.wikipedia.org/wiki/Nuclear_binding_energy#/media/File:Binding_energy_curve_-_common_isotopes.svg, public domain)

Chemistry: Uranium Decay Series



(Image source: https://en.wikipedia.org/wiki/Decay_chain#/media/File:Decay_chain(4n%2B2,_Uranium_series).svg , Tosaka, CC BY 3.0)

Chemistry: Pauli Exclusion Principle, Hund's Rule, Aufbau Principle

Pauli exclusion principle: no two electrons in an atom can have the same four quantum numbers: (n, ℓ, m_ℓ, m_s), an atom's nth electron shell can accommodate $2n^2$ electrons
paramagnetic: refers to substances that are attracted to a magnet
diamagnetic: refers to substances that are slightly repelled by a magnet
Hund's rule: the most stable arrangement of electrons in subshells is the one with the greatest number of parallel spins
Aufbau principle: as protons are added one by one to the nucleus to build up the elements, electrons are similarly added to the atomic orbitals

Chemistry: Rydberg Formula

energy of a photon emitted by a hydrogen atom is given by the difference of two hydrogen energy levels:

$$E = E_i - E_f = R_E\left(\frac{1}{n_f^2} - \frac{1}{n_i^f}\right)$$

R_E = 1.097373 x 10^7 / meter
n_f = final energy level $\quad n_i$ = initial energy level

the wavelength of the emitted photon:

$$\frac{1}{\lambda} = \frac{R_E}{2\pi}\left(\frac{1}{n_f^2} - \frac{1}{n_i^f}\right)$$

Chemistry: Atomic Orbitals (1)

classification according to outermost subshell filled with electrons

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 9 10 ← 8B →	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1s															1s
2s										2p					
3s										3p					
4s		3d								4p					
5s		4d								5p					
6s		5d								6p					
7s		6d													
		4f													
		5f													

Chemistry: Atomic Orbitals (2)

principal quantum number n: describes energy of the electron; always a positive integer; large numbers seldom encountered; each atom has many orbitals associated with each value of n; these orbitals together are sometimes called electron shells
azimuthal quantum number ℓ: describes orbital angular momentum of each electron; a non-negative integer; within a shell where n is some integer n_0, ℓ ranges across all integer values satisfying the relation $0 \le \ell \le n_0 - 1$; set of orbitals associated with a particular value of ℓ are sometimes collectively called a subshell
magnetic quantum number m_ℓ: describes the magnetic moment of an electron in an arbitrary direction; is also always an integer; within a subshell where ℓ is some integer ℓ_0, m_ℓ ranges from $-\ell_0 \le m_\ell \le \ell_0$
electron spin quantum number m_s: counterclockwise +1/2, clockwise -1/2

Chemistry: Atomic Orbitals (3)

	$\ell = 0$	$\ell = 1$	$\ell = 2$	$\ell = 3$	$\ell = 4$	…
n = 1	$m_\ell = 0$					
n = 2	0	-1, 0, 1				
n = 3	0	-1, 0, 1	-2, -1, 0, 1, 2			
n = 4	0	-1, 0, 1	-2, -1, 0, 1, 2	-3, -2, -1, 0, 1, 2, 3		
n = 5	0	-1, 0, 1	-2, -1, 0, 1, 2	-3, -2, -1, 0, 1, 2, 3	-4, -3, -2, -1, 0, 1, 2, 3, 4	
…	…	…	…	…	…	…

Chemistry: Atomic Orbitals (4)

1 H hydrogen $1s^1$	**13 Al aluminum [Ne] $3s^2\ 3p^1$**
2 He helium $1s^2$	**14 Si silicon [Ne] $3s^2\ 3p^2$**
3 Li lithium [He] $2s^1$	**15 P phosphorus [Ne] $3s^2\ 3p^3$**
4 Be beryllium [He] $2s^2$	**16 S sulfur [Ne] $3s^2\ 3p^4$**
5 B boron [He] $2s^2\ 2p^1$	**17 Cl chlorine [Ne] $3s^2\ 3p^5$**
6 C carbon [He] $2s^2\ 2p^2$	**18 Ar argon [Ne] $3s^2\ 3p^6$**
7 N nitrogen [He] $2s^2\ 2p^3$	**19 P potassium [Ar] $4s^1$**
8 O oxygen [He] $2s^2\ 2p^4$	**20 Ca calcium [Ar] $4s^2$**
9 F fluorine [He] $2s^2\ 2p^5$	**21 Sc scandium [Ar] $3d^1\ 4s^2$**
10 Ne neon [He] $2s^2\ 2p^6$	**22 Ti titanium [Ar] $3d^2\ 4s^2$**
11 Na sodium [Ne] $3s^1$	**23 V vanadium [Ar] $3d^3\ 4s^2$**
12 Mg magnesium [Ne] $3s^2$	**24 Cr chromium [Ar] $3d^5\ 4s^1$**

Chemistry: Atomic Orbitals (5)

25 Mn manganese [Ar] $3d^5 4s^2$	37 Rb rubidium [Kr] $5s^1$
26 Fe iron [Ar] $3d^6 4s^2$	38 Sr strontium [Kr] $5s^2$
27 Co cobalt [Ar] $3d^7 4s^2$	39 Y yttrium [Kr] $4d^1 5s^2$
28 Ni nickel [Ar] $3d^8 4s^2$	40 Zr zirconium [Kr] $4d^2 5s^2$
29 Cu copper [Ar] $3d^{10} 4s^1$	41 Nb niobium [Kr] $4d^4 5s^1$
30 Zn zinc [Ar] $3d^{10} 4s^2$	42 Mo molybdenum [Kr] $4d^5 5s^1$
31 Ga gallium [Ar] $3d^{10} 4s^2 4p^1$	43 Tc technetium [Kr] $4d^5 5s^2$
32 Ge germanium] $3d^{10} 4s^2 4p^2$	44 Ru ruthenium [Kr] $4d^7 5s^1$
33 As arsenic [Ar] $3d^{10} 4s^2 4p^3$	45 Rh rhodium [Kr] $4d^8 5s^1$
34 Se selenium [Ar] $3d^{10} 4s^2 4p^4$	46 Pd palladium [Kr] $4d^{10}$
35 Br bromine [Ar] $3d^{10} 4s^2 4p^5$	47 Ag silver [Kr] $4d^{10} 5s^1$
36 Kr krypton [Ar] $3d^{10} 4s^2 4p^6$	48 Cd cadmium [Kr] $4d^{10} 5s^2$

Chemistry: Atomic Orbitals (6)

49 In indium [Kr] $4d^{10} 5s^2 5p^1$	61 Pm promethium [Xe] $4f^5 6s^2$
50 Sn tin [Kr] $4d^{10} 5s^2 5p^2$	62 Sm samarium [Xe] $4f^5 6s^2$
51 Sb antimony [Kr] $4d^{10} 5s^2 5p^3$	63 Eu europium [Xe] $4f^6 6s^2$
52 Te tellurium [Kr] $4d^{10} 5s^2 5p^4$	64 Gd gadolinium [Xe] $4f^7 5d^1 6s^2$
53 I iodine [Kr] $4d^{10} 5s^2 5p^5$	65 Tb terbium [Xe] $4f^9 6s^2$
54 Xe xenon [Kr] $4d^{10} 5s^2 5p^6$	66 Dy dysprosium [Xe] $4f^{10} 6s^2$
55 Cs caesium [Xe] $6s^1$	67 Ha holmium [Xe] $4f^{11} 6s^2$
56 Ba barium [Xe] $6s^2$	68 Er erbium [Xe] $4f^{12} 6s^2$
57 La lanthanum [Xe] $5d^1 6s^1$	69 Tm thulium [Xe] $4f^{13} 6s^2$
58 Ce cerium [Xe] $4f^1 5d^1 6s^2$	70 Yb ytterbium [Xe] $4f^{14} 6s^2$
59 Pr praseodmium [Xe] $4f^3 6s^2$	71 Lu lutetium [Xe] $4f^{14} 5d^1 6s^2$
60 Nd neodymium [Xe] $4f^4 6s^2$	72 Hf hafnium [Xe] $4f^{14} 5d^2 6s^2$

Chemistry: Atomic Orbitals (7)

73 Ta tantalum [Xe] $4f^{14} 5d^3 6s^2$	85 At astatine [Xe] $4f^{14} 5d^{10} 6s^2 6p^5$
74 W tungsten [Xe] $4f^{14} 5d^4 6s^2$	86 Rn radon [Xe] $4f^{14} 5d^{10} 6s^2 6p^6$
75 Re rhenium [Xe] $4f^{14} 5d^5 6s^2$	87 Fr francium [Rn] $7s^1$
76 Os osmium [Xe] $4f^{14} 5d^6 6s^2$	88 Ra radium [Rn] $7s^2$
77 Ir iridium [Xe] $4f^{14} 5d^7 6s^2$	89 Ac actinium [Rn] $6d^1 7s^1$
78 Pt platinum [Xe] $4f^{14} 5d^9 6s^1$	90 Th thorium[Rn] $6d^2 7s^1$
79 Au gold [Xe] $4f^{14} 5d^{10} 6s^1$	91 Pa protactinium [Rn] $5f^2 6d^1 7s^1$
80 Hg mecury [Xe] $4f^{14} 5d^{10} 6s^2$	92 U uranium [Rn] $5f^3 6d^1 7s^1$
81 Tl thallium [Xe] $4f^{14} 5d^{10} 6s^2 6p^1$	93 Np neptunium [Rn] $5f^4 6d^1 7s^1$
82 Pb lead [Xe] $4f^{14} 5d^{10} 6s^2 6p^2$	94 Pu plutonium [Rn] $5f^6 7s^2$
83 Bi bismuth [Xe] $4f^{14} 5d^{10} 6s^2 6p^3$	95 Am americium [Rn] $5f^7 7s^2$
84 Po polonium [Xe] $4f^{14} 5d^{10} 6s^2 6p^4$	96 Cm curium [Rn] $5f^7 6d^1 7s^2$

Chemistry: Atomic Orbitals (8)

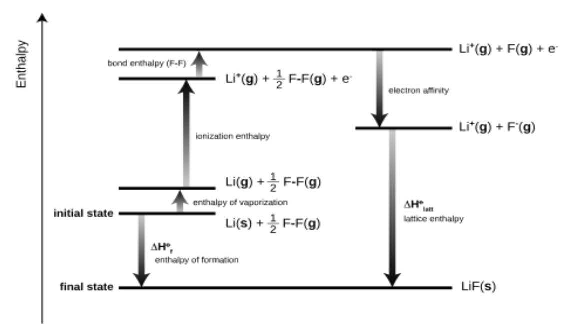
97 Bk berkelium [Rn] $5f^9 7s^2$	109 Mt meitnerium [Rn] $5f^{14} 6d^7 7s^2$
98 Cf californium [Rn] $5f^{10} 7s^2$	110 Ds darmstadtium [Rn] $5f^{14} 6d^8 7s^2$
99 Es einsteinium [Rn] $5f^{11} 7s^2$	111 Rg roentgenium [Rn] $5f^{14} 6d^9 7s^2$
100 Fm fermium [Rn] $5f^{12} 7s^2$	112 Cn copernicium [Rn] $5f^{14} 6d^{10} 7s^2$
101 Md mendelevium [Rn] $5f^{13} 7s^2$	113 Nh nihonium [Rn] $5f^{14} 6d^{10} 7s^2 7p^1$
102 No nobelium [Rn] $5f^{14} 7s^2$	114 Fl flerovium [Rn] $5f^{14} 6d^{10} 7s^2 7p^2$
103 Lr lawrencium [Rn] $5f^{14} 7s^2 7p^1$	115 Mc moscovium [Rn] $5f^{14} 6d^{10} 7s^2 7p^3$
104 Rf rutherfordium [Rn] $5f^{14} 6d^2 7s^2$	116 Lv livermorium [Rn] $5f^{14} 6d^{10} 7s^2 7p^4$
105 Db dubnium [Rn] $5f^{14} 6d^3 7s^2$	117 Ts tennesine [Rn] $5f^{14} 6d^{10} 7s^2 7p^5$
106 Sg seaborgium [Rn] $5f^{14} 6d^4 7s^2$	118 Og oganesson [Rn] $5f^{14} 6d^{10} 7s^2 7p^6$
107 Bh bohrium [Rn] $5f^{14} 6d^5 7s^2$	119 Uue ununennium [Og] $8s^1$
108 Hs hassium [Rn] $5f^{14} 6d^6 7s^2$	120 Ubn unbinilium [Og] $8s^2$

Chemistry: Born-Haber Cycle

Born-Haber cycle: relates lattice energies of ionic compounds to ionization energies, electron affinities, and other atomic and molecular properties

Chemistry: Lattice Energy of Ionic Compounds

lithium-fluorine example



(Image source: https://en.wikipedia.org/wiki/Born%E2%80%93Haber_cycle#/media/File:Born-haber_cycle_LiF.svg, Jkwchui, CC BY-SA 3.0)

Chemistry: Molecular Bonds (1)

bond energy: enthalpy required to break a particular bond in 1 mole of gaseous molecules
ionic bond: electrostatic force that holds ions together in an ionic compound, forms when the electronegativity difference between the two bonding atoms is 2.0 or more
covalent bond: involves the sharing of electron pairs between atoms where the electron pairs are known as shared pairs or bonding pairs
single bond: two atoms are held together by one electron pair
double bond: bonds formed when two atoms share two or more pairs of elctrons
triple bond: when two atoms share three pairs of electrons
resonance: using two or more Lewis structures for a single molecule that cannot be represented accurately by only one Lewis structure
octet rule: an atom other than hydrogen tends to form bonds until it is surrounded by 8 valence electrons

Chemistry: Molecular Bonds (2)

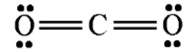
covalent bond: a bond in which two electrons are shared by two atoms
covalent compounds: contain only covalent bonds
polar bond: electrons spend more time in the vicinity of one atom than the other
valence shell: outermost electron-occupied shell of an atom, holds the electrons that are usually involved in bonding
valence-shell electron-pair repulsion (VSEPR) model: accounts for geometric arrangements of electron pairs around a central atom in terms of the electrostatic repulsion between electron pairs
dipole moment: product of charge and distance between the charges
polar molecule: have dipole moments
nonpolar molecules: do not have dipole moments
valence bond theory: stable molecule forms from reacting atoms when the potential energy of the system has decreased to a minimum

Chemistry: Lewis Structure (1)

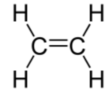
Lewis structure: representation of covalent bonding , shared electron pairs
shown either as lines or as pairs of dots between two atoms, lone pairs are shown as
pairs of dots on individual atoms

(Image source on Lewis Structure (3): https://en.wikipedia.org/wiki/Lewis_structure#/media/File:Infographic_-_Draw_a_Lewis_Dot_Structure._Beaker_Babe_-_2015_(2000x2592).jpg , Lisa Staugaard, CC BY-SA 4.0)

Chemistry: Lewis Structure (2)



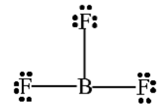
carbon dioxide
double bonds



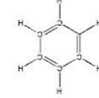
ethelyne
double bonds

nitrogen
molecule
N_2triple bond

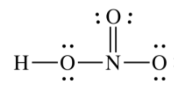
beryllium hydride



boron
hydrofluoride



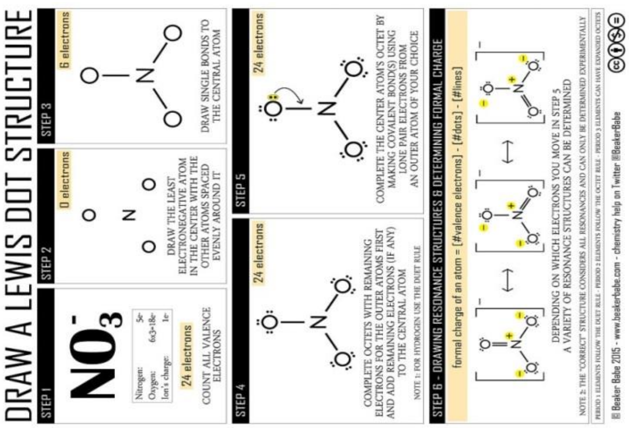
benzene



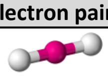
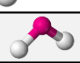
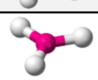
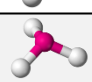
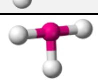
nitric acid

ozone

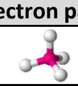
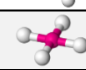
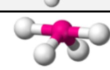
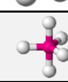
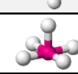
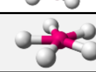
Chemistry: Lewis Structure (3)



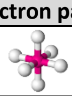
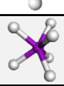
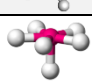
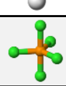
Chemistry: Electron Pair Arrangements around a Central Atom (1)

number of electron pairs	arrangement of electron pairs	geometry	examples
2		linear	BeF_2
2		bent	OF_2
3		trigonal planar	BF_3
3		trigonal pyramidal	NH_3, PCl_3, XeO_3
3		t-shaped	ClF_3

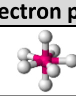
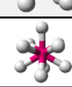
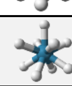
Chemistry: Electron Pair Arrangements around a Central Atom (2)

number of electron pairs	arrangement of electron pairs	geometry	examples
4		tetrahedral	CH_4,PO_4^{3}, SO_2^{4-}
4		square planar	$[Pt(NH_3)_2Cl_2]$
4		seesaw	SF_4
5		trigonal bipyramidal	PF_5
5		square pyramidal	$XeOF_4$
5		pentagonal planar	XeF_5^-

Chemistry: Electron Pair Arrangements around a Central Atom (3)

number of electron pairs	arrangement of electron pairs	geometry	examples
6		octahedral	$SF_6Mo(CO)_6$
6		trigonal prismatic	$W(CH_3)_6$
6		pentagonal pyramidal	$XeOF_5^-$, IOF_5^{2-}
6		distorted octahedral	XeF_6, IF_6^-

Chemistry: Electron Pair Arrangements around a Central Atom (4)

number of electron pairs	arrangement of electron pairs	geometry	examples
7		pentagonal bipyramidal	IF_7
8		square antiprismatic	XeF_8^{2-}
9		tricapped trigonal prismatic	ReH_9^{2-}
9		capped square antiprismatic	$[\{LaCl(H_2O)_7\}_2]^{4+}$

(Image sources: https://en.wikipedia.org/wiki/Linear_molecular_geometry, public domain)

Chemistry: Hybridization

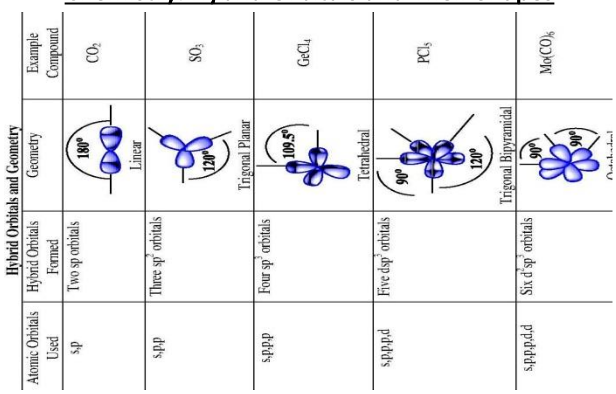
<u>hybrid orbitals</u>: atomic orbitals obtained when two or more nonequivalent orbitals of the same atom combine in preparation for covalent bond formation
<u>hybridization</u>: the mixing of atomic orbitals in an atom to generate a set of hybrid orbitals
<u>sigma bonds</u>: covalent bonds formed by orbitals overlapping end-to-end with the electron density concentrated between the nuclei of the bonding atoms
<u>pi bonds</u>: a covalent bond formed by sideways overlapping orbitals with electron density concentrated above and below the plane of the nuclei of the bonding atoms

Chemistry: Hybridization Rules

- hybridization is used only to explain covalent bonding
- hybridization is the mixing of at least two nonequivalent atomic orbitals; a hybrid is not a pure atomic orbital; hybrid orbitals and pure atomic orbitals have different shapes
- the number of hybrid orbitals generated is equal to the number of pure atomic orbitals that participate in the hybridization
- hybridization requires input energy; the system recovers this energy during bond formation
- covalent bonds in polyatomic molecules and ions are formed by the overlap of hybrid orbitals with unhybridized orbitals

(Source:)

Chemistry: Hybrid Orbitals and Their Shapes



(Image source: Hybrid orbital and geometries by Chem507f10grp4\ , public domain)

Chemistry: Molecular Orbital Theory

<u>molecular orbitals</u>: result from interaction of the atomic orbitals of the bonding atoms and are associated with the entire molecule
<u>bonding molecular orbital</u>: has lower energy and greater stability than the atomic orbitals from which it was formed
<u>antibonding molecular orbital</u>: has higher energy and lower stability than the atomic orbitals from which it was formed
<u>sigma molecular bond</u>: electron density is concentrated symmetrically around a line between the two nuclei of the bonding atoms
<u>bond order</u>: ½(number of electrons in bonding MOs – number of electrons in antibonding MOs)
delocalized molecular orbitals: are not confined between two adjacent bonding atoms, but actually extend over three or more atoms

Chemistry: Molecular Orbital Theory Rules

- number of molecular orbitals formed is always equal to the number of atomic orbitals combined
- the more stable the bonding molecular orbital, the less stable the corresponding antibonding molecular orbit
- filling molecular orbitals proceeds from low to high energies
- each molecular orbital can accommodate up to two electrons with opposite spins, according to the Pauli exclusion principle
- when electrons are added to molecular orbitals of the same energy, the most stable arrangement follows Hund's rule
- number of electrons in the molecular orbitals is equal to the sum of all the electrons on the bonding atoms

(Source:)

Chemistry: Stoichiometry

<u>stoichiometry</u>: quantitative study of reactants and products in a chemical reaction

1. Write correct formulas for reactants and products, balance the equation.
2. Convert the quantities of reactants into moles.
3. Use the coefficients to calculate the number of moles of products.
4. Convert the quantities to grams.
5. Check that the answer is reasonable.

$2H_2$	**$+O_2$**	**→ $2H_2O$**
two moles	**+ one molecule**	**→ two molecules**
2 moles	**+ 1 mole**	**→ 2 moles**
2(2.02 g) = 4.04 g	**+ 32.00 g**	**→ 2(18.02 g) = 36.04 g**

Chemistry: Reaction Rates

<u>elementary steps</u>: series of simple reactions representing progress of overall reaction at the molecular level
<u>reaction mechanism</u>: elementary steps leading to product formation
<u>intermediates</u>: steps that appear in the mechanism of the reaction but not in overall balanced equation
<u>molecularity of a reaction</u>: number of molecules reacting in an elementary step
<u>bimolecular reaction</u>: elementary step that involves two molecules
<u>unimolecular reaction</u>: elementary step in which only one reacting molecule participates
<u>termolecular reaction</u>: reactions involving participation of three molecules in one elementary step
<u>rate determining step</u>: slowest step in the sequence
<u>reaction order</u>: sum of the powers to which all reactant concentrations appearing in the rate law are raised

Chemistry: Equilibrium and Le Chatelier's Principle

homogeneous equilibrium: applies to reactions in which all reacting species are in the same phase

heterogeneous equibrium: results from a reversible reaction involving reactants and products that are in different phases

Le Chatelier's principle: if an external stress is applied to a system at equilibrium the system adjusts so that the stress is partially offset as the system reaches a new equilibrium

change in concentration, pressure, or volume may alter the equilibrium position, but will not change the equilibrium constant; only a change in temperature will change the equilibrium constant;

a catalyst lowers activation energy but does not change the equilibrium constant or shift the position of a system's equilibrium

Chemistry: First and Second Order Reactions

order	rate law	concentration time equation	half-life
1	rate = $\frac{-\Delta[A]}{\Delta t} = k[A]$	$\ln\frac{[A]}{[A]_0} = -kt$	$\frac{0.693}{k}$
2	rate = $\frac{-\Delta[A]}{\Delta t} = k[A]^2$	$\frac{1}{[A]} = \frac{1}{[A]_0} + kt$	$\frac{1}{k[A]_0}$

(Source: p. 520)

Arrhenius equation: graphical determination of activation energy

$$k = Ae^{-E_a/RT} \rightarrow \ln k = \left(\frac{E_a}{R}\right)\left(\frac{1}{T}\right) + \ln A$$

rate constants at two different temperatures $\ln\frac{k_1}{k_2} = \left(\frac{E_a}{R}\right)\left(\frac{T_1 - T_2}{T_1 T_2}\right)$

Chemistry: Equilibrium Constant of a Chemical Reaction (1)

$$aA + bB \Leftrightarrow cC + dD \quad K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b} = \frac{k_{forward}}{k_{reverse}}$$

$$\begin{aligned} A + B &\Leftrightarrow C + D \\ C + D &\Leftrightarrow E + F \end{aligned} \quad K_C{}' = \frac{[C][D]}{[A][B]} \quad K_C{}'' = \frac{[E][F]}{[C][D]} \quad K_C = (K_C{}')(K_c{}'')$$

$$K_P = \frac{P_B^b}{P_A^a} = \frac{[B]^b}{[A]^a}(RT)^{b-a}$$ P_A and P_B are the partial pressures

rate law: if $aA + bB \rightarrow cC + dD$ then rate = $k[A]^x[B]^y$, x and y determined experimentally

Chemistry: Equilibrium Constant of a Chemical Reaction (2)

reaction direction: if $Q_c > K_c$ the ratio of initial concentrations of products to reactants indicates that products must be converted to reactants; if $Q_c < K_c$ the ratio of initial concentrations of products to reactants indicates that reactants must be converted to products; if $Q_c = K_c$ the system is in equilibrium

Q_c = reaction quotient, where initial amounts of product and reactant are substituted into the equation for K_{eq}

Chemistry: Acids and Base

acid: substance that yields hydrogen ions (H^+) when dissolved in water; sour taste; cause color change in plant dyes; react with certain metals, such as zinc, magnesium, and iron to produce hydrogen gas, aqueous acid solutions conduct electricity, pH between 0 and 7

Brönsted acid: proton donor

conjugate acid: results from addition of a proton to a Brönsted base

base: a substance that yields hydroxide ions (OH-) when dissolved in water; bitter taste, feels slippery, cause color change in plant dyes, aqueous base solutions conduct electricity, pH between 7 and 14

Brönsted base: proton acceptor

conjugate base: species that remains when one proton has been removed from the acid

Chemistry: Acids and Bases (2)

neutralization reaction: a reaction between an acid and a base

pH = - log [H_3O^+]   pOH = - log [OH-]   pH + pOH = 14

monoprotic: each unit of acid yields one hydrogen ion upon ionization

diprotic: each unit of acid yields two hydrogen ions in two separate steps

triprotic: each unit of acid yields three hydrogen ions in three separate steps

Chemistry: Acid-Base Properties of Water

$$\underset{acid_1}{H_2O} + \underset{base_2}{H_2O} \Leftrightarrow \underset{acid_2}{H_3O^+} + \underset{base_1}{OH^-} \quad K_C = \frac{[H^+][OH^-]}{[H_2O]}$$

Chemistry: Common Acid-Base Indicators

indicator	in acid	in base	pH range
thymol blue	red	yellow	1.2 – 2.8
bromophenol blue	yellow	bluish purple	3.0 – 4.6
methyl orange	orange	yellow	3.1 – 4.4
methyl red	red	yellow	4.4 – 6.2
methyl purple	purple	green	4.8 – 5.4
phenol red	yellow	red	6.4 – 8.0
naphtholphthalein	pale red	greenish blue	7.3 – 8.7
cresolphthalein	colorless	purple	8.2 – 9.8
phenolphthalein	colorless	purple-pink	8.3 – 10.0
alizarine yellow R	yellow	red	10.2 – 12.0
indigo carmen	blue	yellow	11.4 – 13.0

(Source: https://en.wikipedia.org/wiki/PH_indicator)

Chemistry: Aqueous Solutions

aqueous solution: solute is initially a liquid or a solid and the solvent is water
electrolyte: substance, that when dissolved in water, results in a solution that can conduct electricity
non-electrolyte: substance, that when dissolved in water, does not result in a solution that can conduct electricity
hydration: the process in which an ion is surrounded by water molecules arranged in a specific manner
precipitation reaction: results in the formation of an insoluble product or precipitate
precipitate: insoluble solid that separates from a solution
solubility: the maximum amount of solute that will dissolve in a given quantity of solvent at a specific temperature

Chemistry: Solubility Rules for Common Ionic Compounds in Water at 25^{0}C

soluble compounds	exceptions
compounds containing alkali metal ions (Li^+, Na^+, K^+, Rb^+, Cs^+), ammonium ion (NH_4^+))	
nitrates (NO_3^-),	
bicarbonates (HCO_3^-)	
chlorates (ClO_3^-)	
halides (Cl^-, Br^-, I^-)	halides of Ag^+, Hg_2^{2+}, Pb^{2+}
sulfates (SO_4^{2-})	sulfates of Ag^+, Ca^{2+},Sr^{2+}, Ba^{2+}, Hg^{2+}, Pb^+
insoluble compounds	**exceptions**
carbonates (CO_3^{2-}), phosphates (PO_4^{3-}), chromates (CrO_4^{2-})	compounds containing alkali metal ions and the ammonium ion
sulfides (S^{2-})	
hydroxides (OH^-)	compounds containing alkali metal ions and the Ba^{2+} ion

(Source: p. 109)

Chemistry: Oxidation-Reduction (Redox) Reactions (1)

electrochemistry: branch of chemistry that deals with the relationship between electrical and chemical energy
oxidation reaction: half-reaction that involves loss of electrons; oxidation state increases; oxidation occurs at the anode (positive terminal) in electrolytic reactions
oxidizing agent: substance that accepts electrons and becomes less negative
oxidation number: signifies the number of charges the atom would have in a molecule if the electrons were transferred completely
reduction reaction: half-reaction that involves gain of electrons; oxidation state decreases; reduction occurs at the cathode (negative terminal) in electrolytic reactions
reducing agent: substance that donates electrons and becomes more negative
number of electrons lost by a reducing agent must be equal to the number of electrons gained by an oxidizing agent

Chemistry: Oxidation Reduction (Redox) Reactions (2)

combination reaction: A + B → C
decomposition reaction: C → A + B
displacement reaction: A + BC → AC + B
- hydrogen displacement: all alkali metals and some alkaline earth metals will displace hydrogen from cold water
- metal displacement: a metal in a compound can be displaced by another metal in the element state
- halogen displacement: $F_2 > CL_2 > Br_2 > I_2$, halogens are the most reactive of the nonmetallic elements. fluorine is the strongest known oxidizing agent

disproportion reaction: an element in one oxidation state is simultaneously oxidized and reduced

Chemistry: Calculating Oxidation Number (1)

1. in free elements (in the uncombined state) each atom has an oxidation number of 0
2. for ions composed of only one atom (monoatomic ions) the oxidation number is equal to the charge on the ion; all alkali metals have an oxidation number of +1, all alkaline earth metals have an oxidation number of +2 in their compounds, aluminum has an oxidation number of +3 in all of its compounds
3. the oxidation number of oxygen in most compounds is -2, but in hydrogen peroxide (H_2O_2) and peroxide ion (O_2^{2-}) is -1
4. the oxidation number of hydrogen is +1 except when it is bonded to metals in binary compounds, for example, in LiH, NaH, CaH_2, its oxidation number is -1

Chemistry: Calculating Oxidation Number (2)

5. fluorine has an oxidation number of -1 in all its compounds; other halogens (Cl, Br, I) have negative oxidation numbers when they occur as halide ions in their compounds; when combined with oxygen, for example in oxoacids and oxoanions, they have positive oxidation numbers
6. in a neutral molecule, the sum of the oxidation numbers of all the atoms must be 0; in a polyatomic ion the sum of oxidation numbers of all the elements in the ion must be equal to the net charge of the ion
7. oxidation numbers do not have to be integers; for example, the oxidation number of O in the superoxide ion O_2^-

(Source:)

Chemistry: Types of Solutions

component A	component B	solution state	examples
gas	gas	gas	air
gas	liquid	liquid	CO_2 in water
gas	solid	solid	H_2 gas in palladium
liquid	liquid	liquid	ethanol in water
liquid	solid	liquid	NaCl in water
solid	solid	solid	brass, solder

(Source: p. 468)

Chemistry: Types of Colloids

dispersing medium	dispersed phase	name	example
gas	liquid	aerosol	fog, mist
gas	solid	aerosol	smoke
liquid	gas	foam	whipped cream
liquid	liquid	emulsion	mayonnaise
liquid	solid	sol	milk of magnesia
solid	gas	foam	plastic foam
solid	liquid	gel	jelly, butter
solid	solid	solid sol	alloys, gemstones

(Source: p. 493)

Chemistry: Solution Concentration (1)

solution: homogeneous mixture of two or more substances
dilution: procedure for preparing a less concentrated solution from a more concentrated solution
catalyst: substance that increases the rate of a chemical reaction without being consumed

mole fraction of component A $X_A = \dfrac{moles\ of\ A}{sum\ of\ moles\ of\ all\ components}$

molarity $M = \dfrac{moles\ of\ solute}{1\,liter\ of\ solution}$ molality $m = \dfrac{moles\ of\ solute}{mass\ of\ solvent\ in\ kg}$

$M_i V_i = M_f V_f$ M_i, M_f in molarity V_i, V_f in liters

Chemistry: Solution Concentration (2)

dilution: procedure for preparing a less concentrated solution from a more concentrated solution
titration: a solution of accurately known concentration is gradually added to another solution of unknown concentration until the chemical reaction between the two solutions is complete
equivalence point: point at which an acid has been completely neutralized by a base
indicator: substance that has a distinctly different color in an acid and a base
hydrophilic: water loving
hydrophobic: water fearing

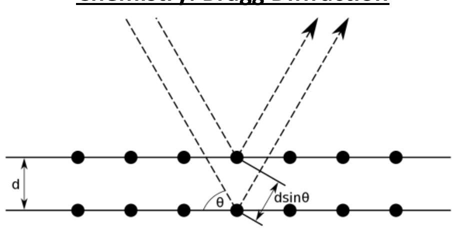
Chemistry: Kinetic and Molecular Theory of Liquids

phase: homogeneous part of the system in contact with other parts of the system but separated by a well-defined boundary
intramolecular forces: hold atoms together in a molecule
van der Waals forces: dipole-dipole, dipole-induced dipole, dispersion forces
dipole-dipole forces: attractive forces between polar molecules
ion-dipole forces: attract an ion and a polar molecule to each other
dispersion forces: attractive forces that arise as a result of temporary dipoles
surface tension: amount of energy required to increase the surface of a liquid by a unit area
cohesion: intermolecular attraction between like molecules
adhesion: an attraction between unlike molecules
viscosity: measure of fluid's resistance to flow
colloid: dispersion of particles of one substance throughout a dispersing medium of another

Chemistry: Kinetic and Molecular Theory of Solids

crystalline solid: possesses rigid and long-range order
unit cell: basic repeating structure of a crystalline solid
coordination number: number of atoms or ions surrounding an atom or ion in a crystal lattice
closest packing: most efficient arrangement of spheres
X-ray diffraction: refers to the scattering of X-rays by the units of a crystalline solid
amorphous solid: lacks a regular three-dimensional arrangement of atoms
glass: an optically transparent fusion product of inorganic materials that has cooled to a rigid state without crystallizing
phase change: transition from one phase to another

Chemistry: Bragg Diffraction



$2d\sin\theta = n\lambda$ d = distance between adjacent planes
θ = angle the X-rays and the plane of the crystal
λ = X-ray wavelength n = integral multiple of wavelength
(Image source: https://en.wikipedia.org/wiki/Bragg's_law#/media/File:BraggPlaneDiffraction.svg, Furiouslettuce, public domain)

Chemistry: Types of Crystals

type	forces	properties	examples
ionic	electrostatic attraction	hard, brittle, high melting point, poor heat and electrical conductor	NaCl, LiF, MgO, $CaCO_3$
covalent	covalent bond	hard, high melting point, poor heat and electrical conductor	C, SiO_2
molecular	dispersion forces, dipole-dipole forces, hydrogen bonds	soft, low melting point, poor heat and electrical conductor	Ar, CO_2, I_2, H_2O, $C_{12}H_{22}O_{11}$
metallic	metallic bond	soft to hard, low to high melting point, good conductor	all metallic elements

Chemistry: Naming Acids and Bases

oxoacid	removal of all H^+ ions →	oxoanion
per - -ic acid	→	per- -ate
↑ +[O]		
representative "–ic" acid	→	-ate
↓ -[O]		
"-ous" acid	→	-ite
↓ -[O]		
hypo- -ous acid	→	hypo- ite

Chemistry: Common Inorganic and Organic Acids

substance	formula	substance	formula
acetic	CH_3COOH	nitric	HNO_3
acrylic	C_2H_3COOH	oleic	$C_{17}H_{33}COOH$
benzene sulfonic	$C_6H_5SO_3H$	oxalic	$H_2C_2O_4$
benzoic	C_6H_5COOH	perchloric	$HClO_4$
butyric	C_3H_7COOH	phenol	C_6H_5OH
carbolic	C_6H_5OH	phosphoric	H_3PO_4
carbonic	H_2CO_3	propionic	C_2H_5COOH
chloric	$HClO_3$	stearic	$C_{17}H_{35}COOH$
formic	HCOOH	sulfuric	H_2SO_4
hydrobromic	HBr	sulfurous	H_2SO_3
hydrochloric	HCl	valeric	C_4H_9COOH
hydrosulfuric	H_2S		

Chemistry: Common Inorganic Anions

bromide (Br^-)	hydroxide (OH^-)
carbonate (CO_3^{2-})	iodide (I^-)
chlorate (ClO_3^-)	nitrate (NO_3^-)
chloride (Cl^-)	nitride (N^{3-})
chromate (CrO_4^{2-})	nitrite (NO_2^-)
cyanide (CN^-)	oxide (O^{2-})
dichromate ($Cr_2O_7^{2-}$)	permanganate (MnO_4^-)
dihydrogen phosphate ($H_2PO_4^-$)	peroxide (O_2^{2-})
fluoride (F^-)	phosphate (PO_4^{3-})
hydride (H^-)	sulfate (SO_4^{2-})
hydrogen carbonate, bicarbonate (HCO_3^-)	sulfide (S^{2-})
hydrogen phosphate (HPO_4^{2-})	sulfite (SO_3^{2-})
hydrogen sulfate, bisulfate (HSO_4^-)	thiocyanate (SCN^-)

Chemistry: Radioactive Decay

radioactive decay $N = N_0 e^{-\lambda t}$ N_0 = original amount

half-life $T_{1/2} = \dfrac{0.693}{\lambda}$ λ = decay constant t = time

daughter product activity

$$N_2 = \frac{\lambda_1 N_{10}}{\lambda_2 - \lambda_1}(e^{-\lambda_1 t} - e^{-\lambda_2 t})$$ λ_1, λ_2 = decay constants

N_{10} = initial activity of parent nuclei

daughter product maximum activity time

$$t' = \frac{\ln \lambda_2 - \ln \lambda_1}{\lambda_2 - \lambda_1}$$

Chemistry: Ideal Gas Law

Ideal gas law: $PV = nRT = NkT$

P = pressure in pascals V = volume in meters^3

n = number of moles T = temperature in kelvins

R = universal gas constant = 8.314 joules/mol K

N = number of molecules

k = Boltzmann's constant = 1.381 x 10^{-23} joules/K = R/N_A

N_A = Avogadro's number = 6.022 x 10^{23}/mol

M = molar mass m = mass in grams

absolute zero = 0 K = -273.15^{0}C

$n = \dfrac{m}{M}$ $\dfrac{n}{V} = \dfrac{m}{MV} = \dfrac{P}{RT}$ density $d = \dfrac{m}{V} = \dfrac{PM}{RT}$

Chemistry: Boyle's Law, Charles' Law, Henry's Law, Raoult's Law

Boyle's Law: $P_i V_i = P_f V_f$ constant temperature

Charles' Law: $\dfrac{V_i}{T_i} = \dfrac{V_f}{T_f}$ constant pressure

Henry's Law: the solubility of a gas in a liquid is proportional to the pressure of the gas over the solution $c = kP$

c = molar concentration in mol/liter

k in units of mol/liter atm P = pressure of the gas over the solution

Raoult's Law: $P_1 = X_1 P_1^0$ partial pressure of a solvent over a solution P_1 is the vapor pressure of the pure solvent P_1^0 times the mole fraction of the solvent in solution X_1

Chemistry: Kinetic Theory of Ideal Gases

average kinetic energy of an ideal gas: $K_{avg} = \dfrac{3}{2}kT$

root mean square velocity: $v_{rms} = \sqrt{\dfrac{3RT}{M}}$

diffusion: the gradual mixing of molecules of one gas with molecules of another by virtue of their kinetic properties

Chemistry: Common Gases

elements	compounds
H_2 (molecular hydrogen)	HF (hydrogen fluoride)
N_2 (molecular nitrogen)	HCl (hydrogen chloride)
O_2 (molecular oxygen)	HBr (hydrogen bromide)
O_3 (ozone)	HI (hydrogen iodide)
F_2 (molecular fluorine)	CO (carbon monoxide)
Cl_2 (molecular chlorine)	CO_2 (carbon dioxide)
He (helium)	NH_3 (ammonia)
Ne (neon)	NO (nitric oxide)
Ar (Argon)	NO_2 (nitrogen dioxide)
Kr (krypton)	N_2O (nitrous oxide)
Xe (xenon)	SO_2 (sulfur dioxide)
Rn (radon)	H_2S (hydrogen sulfide)
	HCN (hydrogen cyanide)

Chemistry: Partial Pressure

Dalton's law of partial pressure: the total pressure of a mixture of gases is the sum of the pressures that each gas would exert if it were present alone

$$P_A = \frac{n_A RT}{V} \qquad P_B = \frac{n_B RT}{V}$$

P_A = pressure exerted by gas *A*

n_A = number of moles of gas *A*

P_B = pressure exerted by gas *B*

n_B = number of moles of gas *B*

total pressure $P_T = P_A + P_B$

Chemistry: Atmospheric Partial Pressure

partial pressure of oxygen:

$$P_{O_2} = X_{O_2} P_{total} = \frac{n_{O_2}}{n_{O_2} + n_{N_2}} P_{total} = \frac{V_{O_2}}{V_{O_2} + V_{N_2}} P_{total}$$

$$= \frac{\sim 20\%}{\sim 20\% + \sim 80\%}(1\,atm) = 0.20\,atm$$

to maintain the same partial oxygen pressure of P_{total} = 0.20 atm:

$$P_{O_2} = \frac{\sim 10\%}{\sim 10\% + \sim 90\%}(2\,atm) = 0.20\,atm$$

partial pressure of nitrogen:

$$P_{N_2} = X_{N_2} P_{total} = \frac{n_{N_2}}{n_{O_2} + n_{N_2}} P_{total} = \frac{V_{N_2}}{V_{O_2} + V_{N_2}} P_{total}$$

Chemistry: van der Waals Equation

van der Waals equation $\left(P+\frac{an^2}{V^2}\right)(V-nb)=nRT$

gas	*a* (atm L^2/mol^2)	*b* (L/mol)	gas	*a* (atm L^2/mol^2)	*b* (L/mol)
He	**0.034**	**0.0237**	**O_2**	**1.36**	**0.0318**
Ne	**0.211**	**0.0171**	**Cl_2**	**6.49**	**0.0562**
Ar	**1.34**	**0.0322**	**CO_2**	**3.59**	**0.0427**
Kr	**2.32**	**0.0398**	**CH_4**	**2.25**	**0.0428**
Xe	**4.19**	**0.0266**	**CCl_4**	**20.4**	**0.138**
H_2	**0.244**	**0.0266**	**NH_3**	**4.17**	**0.0371**
N_2	**1.39**	**0.0391**	**H_2O**	**5.46**	**0.0305**

Chemistry: Energy Changes

thermochemistry: study of heat change in chemical reactions
open system: can exchange mass and energy, usually in the form of heat with the surroundings
closed system: allows the transfer of energy, usually in the form of heat, but not mass, with the surroundings
isolated system: does not allow the transfer of mass or energy
exothermic process: any process that emits heat to the surroundings
endothermic process: any process that requires heat from the surroundings
change of enthalpy: represents the amount of heat given off or absorbed during a reaction
heat capacity: mass x specific heat
standard enthalpy of formation: heat change resulting when one mole of a compound is formed from its elements at a pressure of 1 atm

Chemistry: Types of Reactions Involving Heat

heat of neutralization: when one equivalent of an acid and one equivalent of a base undergo a neutralization reaction to form water and a salt; defined as the energy released with the formation of 1 mole of water

heat of ionization: increase in enthalpy when 1 mole of a substance is completely ionized at constant pressure

heat of fusion: change in its enthalpy resulting from providing energy, typically heat, to a specific quantity of the substance to change its state from a solid to a liquid at constant pressure

Chemistry: Types of Reactions Involving Heat (2)

heat of vaporization: energy (enthalpy) that must be added to the liquid substance, to transform a quantity of that substance into a gas. The enthalpy of vaporization is a function of the pressure at which that transformation takes place

heat of reaction: change in the enthalpy of a chemical reaction that occurs at a constant pressure

heat of solution: heat generated or absorbed when a certain amount of solute dissolves in a certain amount of solvent

heat of dilution: heat change associated with the dilution process

Chemistry: Calorimetry and Latent Heat

for an isolated system

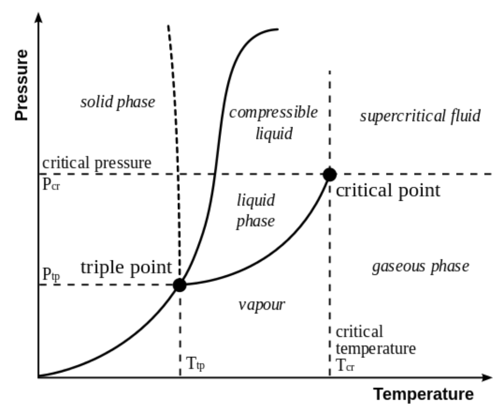
$$Q_{system}=Q_{water}+Q_{cal}+Q_{reaction}=0$$
$$Q_{water}=m_{water}c_{water}\Delta T_{water}$$
$$Q_{cal}=m_{cal}c_{cal}\Delta T_{cal}$$
$$Q_{reaction}=m_{reaction}c_{reaction}\Delta T_{reaction}$$

latent heat of fusion $L_{fusion}=\frac{Q_{fusion}}{m_{material}}$

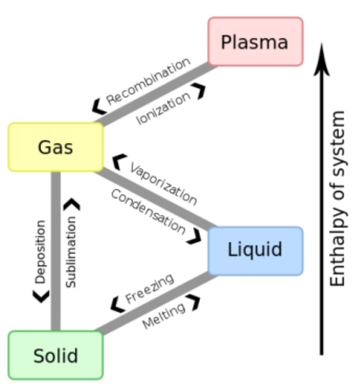
latent heat of vaporization $L_{vaporization}=\frac{Q_{vaporization}}{m_{material}}$

Chemistry: Phase Change Diagram

(Image source: https://en.wikipedia.org/wiki/Phase_transition#/media/File:Phase-diag2.svg , Matthieumarechal, CC BY-SA 3.0)



Physics Phase Changes (1)



(Image source: https://commons.wikimedia.org/wiki/File:Phase_change_-_en.svg , Flanker, penubag, public domain)

Chemistry: Phase Changes (2)

		to			
		solid	**liquid**	**gas**	**plasma**
from	**solid**	**--**	**melting**	**sublimation**	**--**
	liquid	**freezing**	**--**	**boiling/ evaporation**	**--**
	gas	**deposition**	**condensation**	**--**	**ionization**
	plasma	**--**	**--**	**recombination/ deionization**	**--**

Chemistry: Heat of Vaporization and Clausius-Clapeyron Equation

$$\ln P = \frac{-\Delta H_{vap}}{RT} + C$$

P = vapor pressure
ΔH_{vap} = molar heat of vaporization required to vaporize 1 mole of a liquid
R = gas constant, 8.314 J/K mol
T = temperature

Chemistry: Critical Temperature and Pressure

critical temperature: temperature above which a substance's gas phase cannot be made to liquify no matter how great the pressure is

critical pressure: minimum pressure that must be applied to bring about liquification at the critical temperature

Chemistry: Batteries (1)

battery: an electrochemical cell, or a series of combined electrochemical cells, that can be used as a source of direct electric current at a constant voltage

dry cell battery: anode consists of a zinc container in contact with manganese dioxide (MnO_2) and an electrolyte, consisting of ammonium chloride and zinc chloride in water with added starch
anode: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
cathode: $2NH_4^+(aq) + 2MnO_2(s) + 2e^- \rightarrow MnO_3(s) + 2NH_3(aq) + H_2O(l)$
overall: $Zn(s) + 2NH_4^+ + 2MnO_2(s) \rightarrow$
$Zn^{2+}(aq) + 2NH_3(aq) + H_2O(l) + Mn_2O_3(s)$

Chemistry: Batteries (2)

mercury battery: contained in a stainless steel cylinder, consists of a zinc anode amalamated with mercury in contact with a strongly alkaline electrolyte containing zinc oxide and mercury(II) oxide
anode: $Zn(Hg) + 2OH^-(aq) \rightarrow ZnO(s) + H_2O(l) + 2e^-$
cathode: $HgO(s) + H_2O(l) + 2e^- \rightarrow Hg(l) + 2OH^-(aq)$
overall: $Zn(Hg) + HgO(s) \rightarrow ZnO(s) + Hg(l)$

solid-lithium battery: employs a solid as the electrolyte connecting the electrodes, cathode made of either TiS_2 or V_6O_{13}, can be recharged

Chemistry: Batteries (3)

lead storage battery: each cell has a lead anode and a cathode of lead dioxide (PbO_2) packed on a metal plate
anode: $Pb(s) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) + 2e^-$
cathode: $PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O(l)$
overall: $Pb(s) + PbO_2(s) + 4H^+(aq) + 2SO_4^{2-}(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$

rechargeable, replenishing the original materials:
$PbSO_4(s) + 2e^- \rightarrow Pb(s) + SO_4^{2-}(aq)$
$PbSO_4(s) + 2H_2O(l) \rightarrow PbO_2(s) + 4H^-(aq) + SO_4^{2-}(aq) + 2e^-$
overall: $PbSO_4(s) + 2H_2O(l) \rightarrow Pb(s) + PbO_2(s) + 4H^+(aq) + 2SO_4^{2-}(aq)$

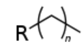
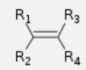
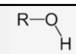
Chemistry: Fuel Cells and Electrolysis in Water

fuel cell: an electrochemical cell that requires a continuous supply of reactants to keep it functioning
hydrogen-oxygen fuel cell: consists of an electrolyte solution, such as potassium hydroxide solution and two inert electrodes, hydrogen and oxygen are bubbled through the anode and cathode compartments
anode: $2H_2(g) + 4OH^-(aq) \rightarrow 4H_2O(l) + 4e^-$
cathode: $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$
overall: $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$
electrolysis in water in a 0.1 molar H_2SO_4 solution:
anode: $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$
cathode: $4\,[\,H^+(aq) + e^- \rightarrow H_2(g)/2\,]$
overall: $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$

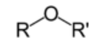
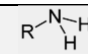
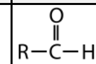
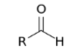
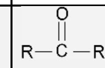
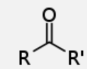
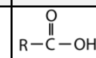
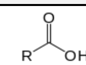
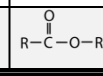
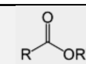
Chemistry: Formation of Organic Compounds

organic chemistry: the branch of chemistry that deals with carbon compounds
aliphaic hydrocarbons: do not contain the benzene group or the benzene ring
aromatic hydrocarbons: contain one or more benzene rings
oxidation: replacement of a hydrogen atom with a hydroxyl group
reduction: replacement of a hydroxyl group with a hydrogen atom
hydrogenation: addition of hydrogen atoms
hydrolysis: addition of one or more water molecules
dehydration: removal of one or more water molecules
polymer: molecular compound distinguished by a high molar mass, ranging into thousands and millions of grams, and made up of many repeating units

Chemistry: Organic Compounds (1)

	IUPAC name	common name	general formula	structural formula
alkane	ethane	ethane	RH	
alkene	ethene/ ethylene	ethylene	$RCH{=}CH_2$ $RCH{=}CHR$ $R_2C{=}CHR$ $R_2C{=}CR_2$	
alkyne	ethyne/ acetylene	acetylene	RC ≡ CH RC ≡ CR	RC≡C-R'
arene	benzene	benzene	ArH	ArH
haloalkane	chloro-ethane	ethyl chloride	RX	R-X
alcohol	ethanol	ethyl alcohol	ROH	

Chemistry: Organic Compounds (2)

	IUPAC name	common name	general formula	structural formula
ether	methoxy-mexane	dimethyl ether	ROR	
amine	methan-amine	methyl-amine	RNH_2 R_2NH, R_3	
aldehyde	ethanal	acetal-dehyde		
ketone	acetone	dimethyl ketone		
carboxylic acid	ethanoic acid	acetic acid		
ester	methyl ethanoate	methyl acetate		

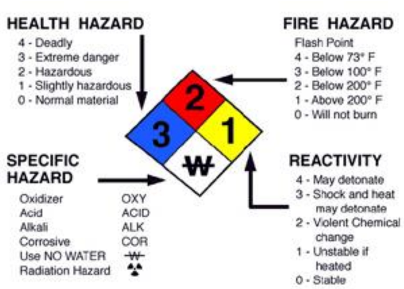
Chemistry: Common Names of Organic and Inorganic Chemicals (1)

common name	chemical name	formula
muriatic acid	hydrochloric acid	HCl
cumene	isopropyl benzene	$C_6H_5CH(CH_3)_2$
styrene	vinyl benzene	$C_6H_5CH{=}CH_2$
epsom salt	magnesium sulfate	$MgSO_4$
hydroquinone	p-dihydroxy benzene	$C_6H_4(OH)_2$
soda ash	sodium carbonate	Na_2CO_3
potash	potassium carbonate	K_2CO_3
baking soda	sodium bicarbonate	$NaHCO_3$
lye/caustic soda	sodium hydroxide	$NaOH$
carbolic acid	phenol	C_6H_5OH

Chemistry: Common Names of Organic and Inorganic Chemicals (2)

common name	chemical name	formula
aniline	aminobenzene	$C_6H_5NH_2$
toluene	methyl benzene	$C_6H_5CH_3$
xylene	dimethyl benzene	$C_6H_4(CH_3)_2$
neopentane	2,2-dimethylpropane	$CH_3C(CH_3)_2CH_3$
heavy water	deuterium oxide	$(H^2)_2O$
eyewash	boric acid solution	H_3BO_3
laughing gas	nitrous oxide	N_2O
brine	sodium chloride solution	$NaCl$
battery acid	sulfuric acid	H_2SO_4

Chemistry: Fire Hazard Diamond (1)



(Image source: https://riskmanagement.unt.edu/riskman/index.php?section=onlinetraining&group=artsafety&module=3)

Chemistry: Fire Hazard Diamond (2)

Position A (left, blue) – Hazard

0 = ordinary combustible hazard
1 = slightly hazardous
2 = hazardous
3 = extreme danger
4 = deadly

Position B (top, red) – Flammability

0 = will not burn
1 = will ignite if preheated
2 = will ignite if moderately heated
3 = will ignite at most ambient temperature
4 = burns readily at ambient conditions 100°F (38°C).

Chemistry: Fire Hazard Diamond (3)

Position C (right, yellow) – Reactivity

0 = stable and not reactive with water
1 = unstable if heated
2 = violent chemical change
3 = shock short may detonate
4 = may detonate

Position D (bottom, white)

OXY = oxidizer ACID = acid ALKALI = alkali

Cor = corrosive W = use no water ☢ = radiation hazard

Flammable

Describes any solid, liquid, vapor, or gas that will ignite easily and burn rapidly. A flammable liquid is defined by NFPA and DOT as a liquid with a flash point below 100^{0}F (38^{0}C)