

Molecules, ions and their compounds.



Antoine Lavoisier
(1743-1794)

TRAITÉ
ÉLÉMENTAIRE
DE CHIMIE,
PRÉSENTÉ DANS UN ORDRE NOUVEAU
ET D'APRÈS LES DÉCOUVERTES MODERNES;

Avec Figures :

Par M. LAVOISIER, de l'Académie des
Sciences, de la Société Royale de Médecine, des
Sociétés d'Agriculture de Paris & d'Orléans, de
la Société Royale de Londres, de l'Institut de
Bologne, de la Société Helvétique de Basle, de
celles de Philadelphie, Harlem, Manchester,
Padoue, &c.

TOME PREMIER.



A PARIS,

Chez CUCHET, Libraire, rue & hôtel Serpente.

M. DCC LXXXIX.

Sous le Privilège de l'Académie des Sciences & de la
Société Royale de Médecine.

Chemical formula

There are different ways of representing a chemical formula

Name	Molecular formula	Extended formula	Structural formula	Molecular model
Ethanol	C_2H_6O	CH_3CH_2OH		
Di-methyl ether	C_2H_6O	CH_3OCH_3		

Colour code for atom types in molecular models



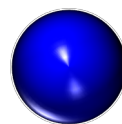
C



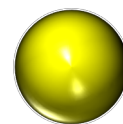
H



O



N



S

Inorganic chemistry compounds

Oxide: compounds containing oxygen

- oxides
- peroxides
- superoxides
- hydroxides (OH^-)

Acids:

- oxides + water (oxoacids)
- elements + hydrogen (hydrides)

Salts: acid + hydroxide

Coordination compound:

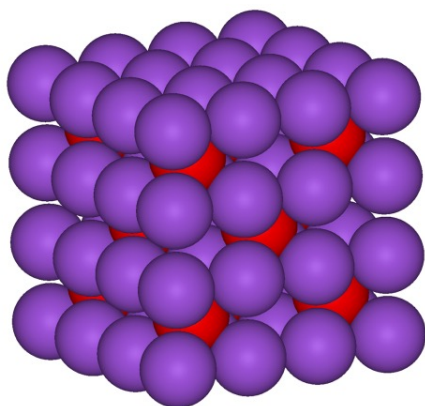
Compound in which an element is bound with more atoms than its o.n. would allow.

Oxidation number

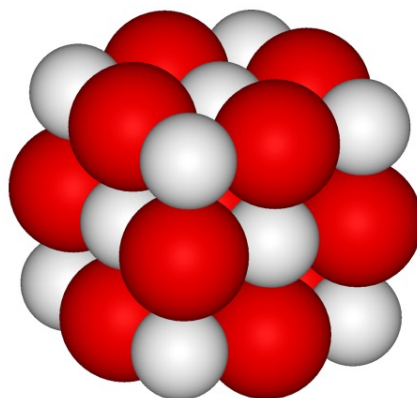
1. In free elements (that is, in uncombined state), each atom has an oxidation number of zero. Na , Fe , C , H_2 , Cl_2 , O_2 , etc.
2. In a monoatomic ion, it corresponds to the charge.
 Na^+ , Ba^{2+} , Fe^{3+} , Br^- , S^{2-} o.n. +1, +2, +3, -1, -2
3. Hydrogen has always oxidation number +1, with the exception of metal hydrides -1
o.n. $\text{H} = +1$, HCl , H_2O , HNO_3 , NH_3 , NH_4^+ , etc.
o.n. $\text{H} = -1$ NaH , KH , CaH_2
4. Oxygen has oxidation number -2 in most compound, with the exception of peroxides, compounds with F and superoxides.
o.n. $\text{O} = -2$, H_2O , BaO , H_2SO_4 , Cl_2O_5 , etc.
o.n. $\text{O} = -1$, H_2O_2 , Na_2O_2 , etc.
o.n. $\text{O} = +2$, OF_2
o.n. = -0.5 NaO_2 ,
5. The algebraic sum of o.n. in an uncharged compound must be 0
in NaCl , o.n. $\text{Na} = +1$ and o.n. $\text{Cl} = -1$
6. The algebraic sum of o.n. in a polyatomic ion must be the charge of the ion. (NH_4^+ , SO_4^{2-} , PO_4^{3-}) PO_4^{3-} : o.n. $\text{O} = -2$ and o.n. $\text{P} = +5$

Oxides of metals

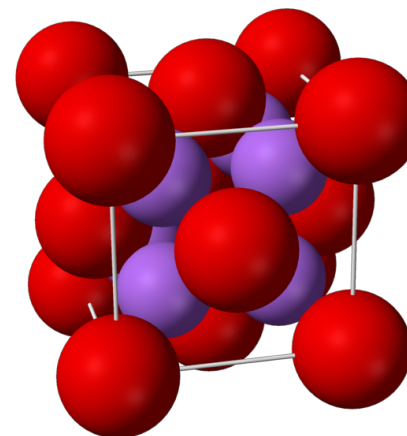
Oxides of elements of Groups I and II are **ionic compounds**. The metal is a cation and oxygen is in the anionic form O^{2-} .



K_2O



CaO



Na_2O



$2 K^+$ and $1 O^{2-}$



$1 Ca^{2+}$ and $1 O^{2-}$



$2 Na^+$ and $1 O^{2-}$

Most oxides of metal are **basic**. They react with water yielding **hydroxides**.

$\text{Na}_2\text{O} + \text{H}_2\text{O} \rightarrow 2 \text{NaOH}$	Sodium hydroxide
$\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2$	Calcium dihydroxide
$\text{FeO} + \text{H}_2\text{O} \rightarrow \text{Fe(OH)}_2$	Iron dihydroxide
$\text{Fe}_2\text{O}_3 + 3 \text{H}_2\text{O} \rightarrow 2 \text{Fe(OH)}_3$	Iron trihydroxide
$\text{Al}_2\text{O}_3 + 3 \text{H}_2\text{O} \rightarrow 2 \text{Al(OH)}_3$	Aluminum trihydroxide

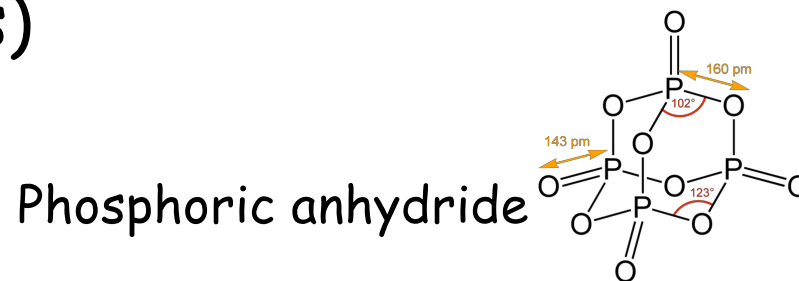
Hydroxides

Ternary complexes formed by H, O and a metal, ionic compounds.
Formula for hydroxides: XOH (metal - hydroxide ions).

Hydroxides have **basic** properties. Hydroxides from alkali metals (IA group) are **strong bases**.

hydroxide	name
NaOH	Sodium Hydroxide
KOH	Potassium Hydroxide
Mg(OH) ₂	Magnesium dihydroxide
Ca(OH) ₂	Calcium dihydroxide
Zn(OH) ₂	zinc dihydroxide
Fe(OH) ₂	Iron dihydroxide (ferrous)
Al(OH) ₃	Aluminun trihydroxide
Cr(OH) ₃	Chrome trihydroxide
Fe(OH) ₃	iron trihydroxide (ferric)

Oxides of nonmetals (anhydrides)



In oxides oxygen always has **oxidation number = -2**.

Oxides with non metals are also called **anhydrides**:

CO_2 (carbon dioxide)

SO_2 (sulphur dioxide)

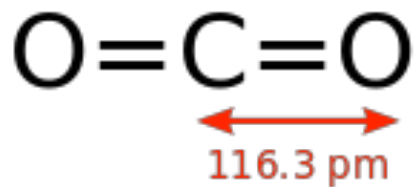
SO_3 (sulphur trioxide)

IUPAC has suffixes **mono-, di-, tri-, tetra-, penta-** to indicate the number of O atoms.

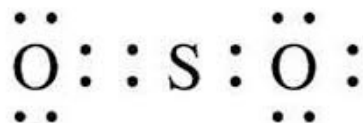
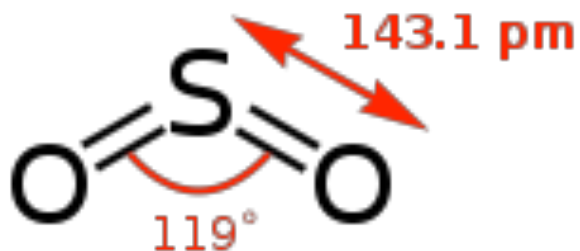
NO : nitrogen monoxide (nitric oxide)

NO_2 : nitrogen dioxide

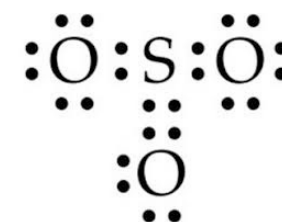
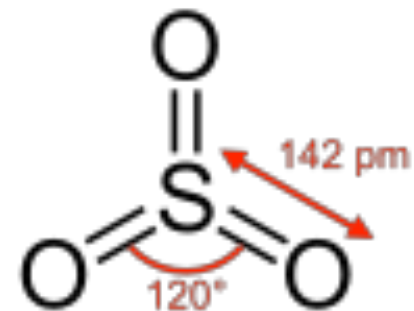
N_2O_3 : dinitrogen trioxide



Carbon dioxide
(carbon anhydride)

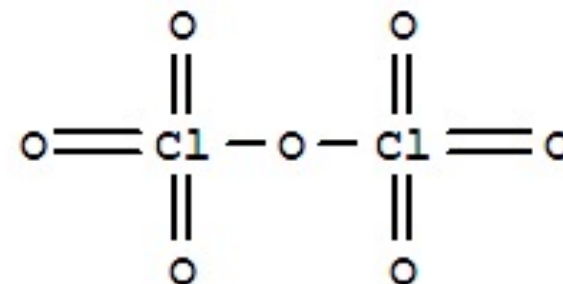
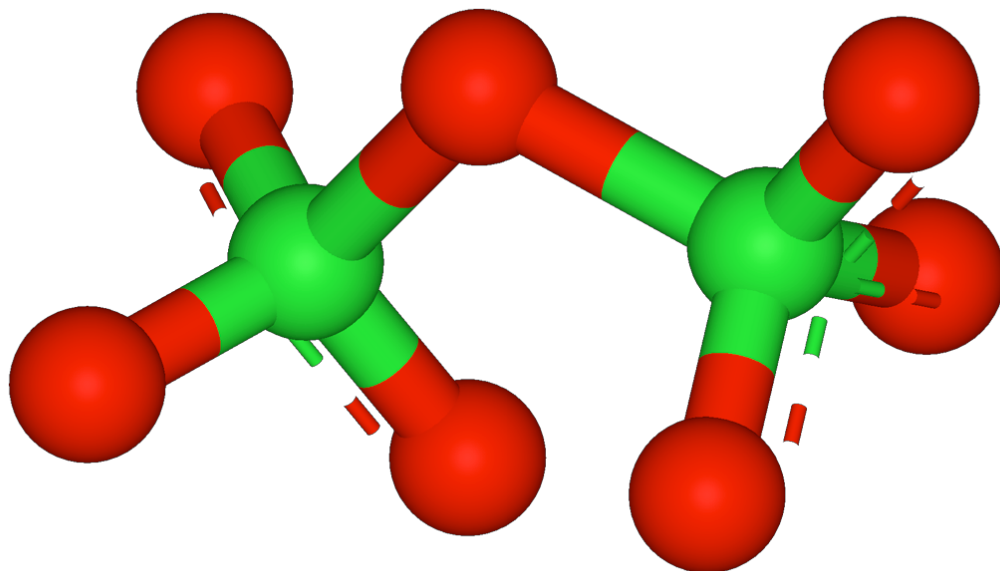


Sulfur dioxide
(sulfurous anhydride o.n.=-4)



Sulfur trioxide
(sulfuric anhydride o.n.=-6)

NB: suffix **-ous** = **lower** oxidation number
 suffix **-ic** = **higher** oxidation number



Dichlorine heptoxide

Although it is the most stable chlorine oxide, Cl_2O_7 is a strong oxidizer as well as an explosive that can be set off with flame or mechanical shock, or by contact with iodine. Nevertheless, it is less strongly oxidising than the other chlorine oxides, and does not attack sulfur, phosphorus, or paper when cold. It has the same effects on the human body as elemental chlorine, and requires the same precautions.

Most oxides of **nonmetals** have acid properties. They react with water yielding **oxoacids (or oxyacids)**.

$\text{N}_2\text{O}_3 + \text{H}_2\text{O} \rightarrow 2 \text{HNO}_2$	Nitrous acid
$\text{N}_2\text{O}_5 + \text{H}_2\text{O} \rightarrow 2 \text{HNO}_3$	Nitric acid
$\text{SO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_3$	Sulphurous acid
$\text{SO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_4$	Sulphuric acid
$\text{Cl}_2\text{O} + \text{H}_2\text{O} \rightarrow 2 \text{HClO}$	Hypochlorous acid
$\text{Cl}_2\text{O}_3 + \text{H}_2\text{O} \rightarrow 2 \text{HClO}_2$	Chlorous acid
$\text{Cl}_2\text{O}_5 + \text{H}_2\text{O} \rightarrow 2 \text{HClO}_3$	Chloric acid
$\text{Cl}_2\text{O}_7 + \text{H}_2\text{O} \rightarrow 2 \text{HClO}_4$	Perchloric acid
$\text{CO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{CO}_3$	Carbonic acid
$\text{P}_2\text{O}_3 + 3 \text{H}_2\text{O} \rightarrow 2 \text{H}_3\text{PO}_3$	Phosphorous acid
$\text{P}_2\text{O}_5 + 3 \text{H}_2\text{O} \rightarrow 2 \text{H}_3\text{PO}_4$	Phosphoric acid

NB: suffix **-ous** = **lower** oxidation number
suffix **-ic** = **higher** oxidation number

Some common oxides (O₂ Oxidation number= −2)

	compound	IUPAC name	Common name (obsolete)	Oxidation number
metals	Na ₂ O	Disodium oxide	-	+1
	K ₂ O	Dipotassium oxide	-	+1
	MgO	Magnesium oxide	-	+2
	CaO	Calcium oxide	-	+2
	Al ₂ O ₃	Dialuminum trioxide	-	+3
	FeO	Iron (II) oxide	Ferrous oxide	+2
	Fe ₂ O ₃	Iron (III) oxide	Ferric oxide	+3
non metals	CO	Carbon monoxide		+2
	CO ₂	Carbon dioxide	Carbonic anhydride	+4
	N ₂ O	Dinitrogen oxide	Nitrogen protoxide	+1
	NO	Nitrogen monoxide	-	+2
	N ₂ O ₃	Trinitrogen dioxide		+3
	NO ₂	Nitrogen dioxide	-	+4
	N ₂ O ₅	dinitrogen pentoxide	Nitric anhydride	+5
	SO ₂	Sulphur dioxide	Sulphurous anhydride	+4
	SO ₃	Sulphur trioxide	Sulphuric anhydride	+6
	Cl ₂ O	Dichlorine oxide	Hypochlorous anhydride	+1
	Cl ₂ O ₃	Dichlorine trioxide	Chlorous anhydride	+3
	Cl ₂ O ₅	Dichlorine pentoxide	Chloric anhydride	+5
	Cl ₂ O ₇	Dichlorine heptoxide	Perchloric anhydride	+7

Peroxides

Compounds containing the group: $-O - O -$
Oxygen has oxidation number = -1

They can be **covalent** (hydrogen peroxide, H_2O_2)
or **ionic** (sodium peroxide, Na_2O_2 , calcium peroxide, CaO_2) peroxidic
group: O_2^{2-}

Superoxides

Ionic compounds containing the ion: O_2^-
In superoxide, it has oxidation number $-1/2$

Combination of the metal ion with the superoxid anion (O_2^-),
es. NaO_2 , KO_2

SUPEROXIDE ION

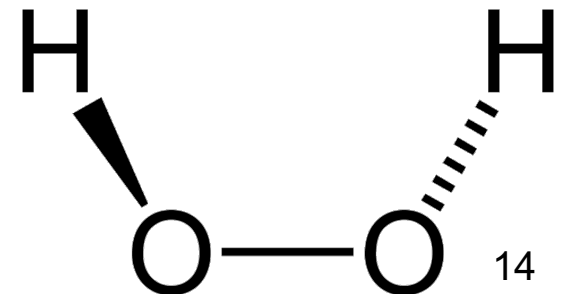


The six outer electrons of each atom of oxygen are highlighted in black; an electron pair is shared; The unpaired electron is shown in the upper left and the extra electron which gives the negative charge is shown in red.

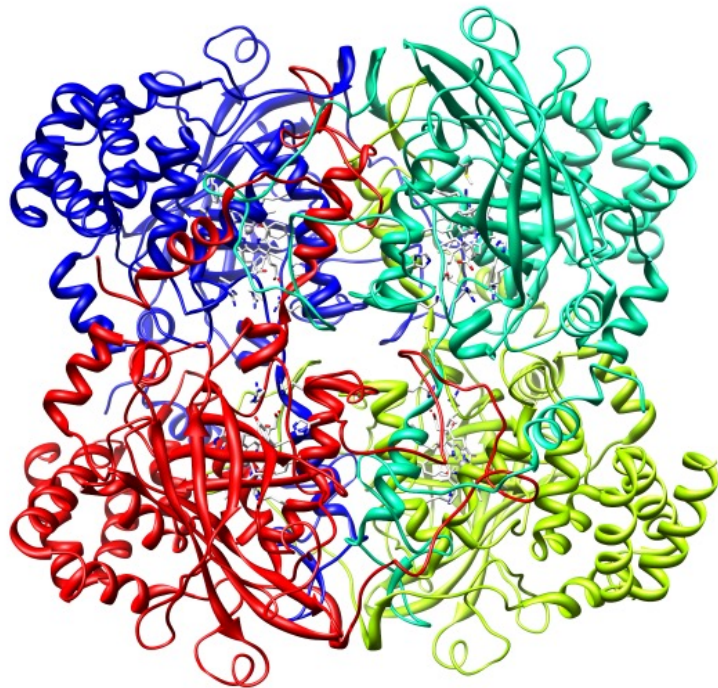
Superoxide is toxic and is used by the immune system as defense against pathogenic microorganisms.

Peroxides are compounds containing this consisting of two oxygen atoms joined by a single bond (O-O).

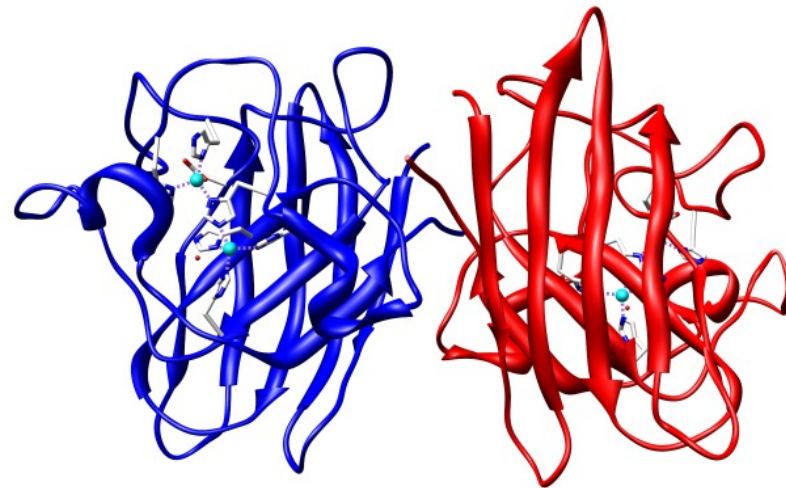
Hydrogen peroxide



Hydrogen peroxide and superoxide ion are oxidizing agents, toxic for the cell. Two enzymes can transform these compounds into less harmful products.



Human catalase



Bovine superoxide dismutase.

The peroxisomal disorder acatalasia is due to a deficiency in the function of catalase. Genetic polymorphisms in catalase and its altered expression and activity are associated with oxidative DNA damage and subsequently the individual's risk of cancer susceptibility

Mice lacking SOD2 die several days after birth, amid massive oxidative stress. Mice lacking SOD1 develop a wide range of pathologies, including hepatocellular carcinoma, an acceleration of age-related muscle mass loss, an earlier incidence of cataracts and a reduced lifespan. Mice lacking SOD3 do not show any obvious defects and exhibit a normal lifespan, though they are more sensitive to hyperoxic injury.

Electronegativity determines the polarization of the chemical bond

- Metals donate electrons yielding cations
- Nonmetals acquire electrons yielding anions.

1A																	8A
1 H	2A											3A	4A	5A	6A	7A	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 –	113 –	114 –	115 –	116 –		

metals

Transition metals

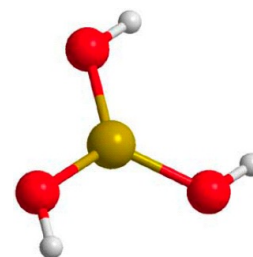
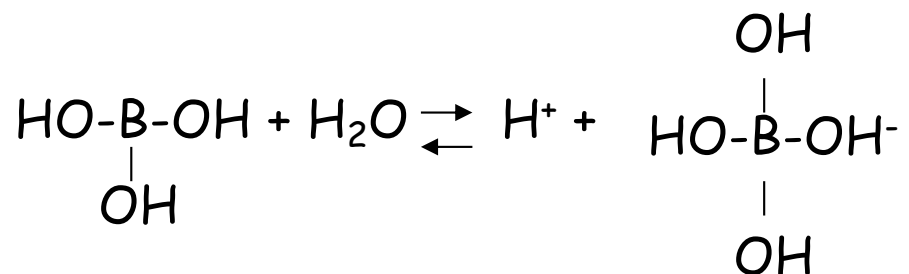
metalloids

Nonmetals
17

Elements of group III: B vs Al

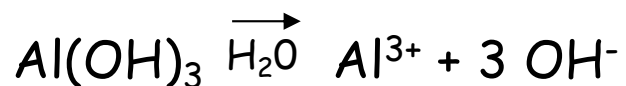
H_3BO_3 Boric acid.

Reacts with water:



$\text{Al}(\text{OH})_3$ Aluminum oxide

Reacts with water:



The difference in behaviour depends on difference in polarization of the bonds B-O e Al-O

Most oxides of **nonmetals** have acid properties. They react with water yielding **oxoacids (or oxyacids)**.

$\text{N}_2\text{O}_3 + \text{H}_2\text{O} \rightarrow 2 \text{HNO}_2$	Nitrous acid
$\text{N}_2\text{O}_5 + \text{H}_2\text{O} \rightarrow 2 \text{HNO}_3$	Nitric acid
$\text{SO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_3$	Sulphurous acid
$\text{SO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_4$	Sulphuric acid
$\text{Cl}_2\text{O} + \text{H}_2\text{O} \rightarrow 2 \text{HClO}$	Hypochlorous acid
$\text{Cl}_2\text{O}_3 + \text{H}_2\text{O} \rightarrow 2 \text{HClO}_2$	Chlorous acid
$\text{Cl}_2\text{O}_5 + \text{H}_2\text{O} \rightarrow 2 \text{HClO}_3$	Chloric acid
$\text{Cl}_2\text{O}_7 + \text{H}_2\text{O} \rightarrow 2 \text{HClO}_4$	Perchloric acid
$\text{CO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{CO}_3$	Carbonic acid
$\text{P}_2\text{O}_3 + 3 \text{H}_2\text{O} \rightarrow 2 \text{H}_3\text{PO}_3$	Phosphorous acid
$\text{P}_2\text{O}_5 + 3 \text{H}_2\text{O} \rightarrow 2 \text{H}_3\text{PO}_4$	Phosphoric acid

NB: suffix **-ous** = **lower** oxidation number
suffix **-ic** = **higher** oxidation number

Oxyacids

Ternary complexes formed by **H**, **O** and a **nonmetal**, or a metal whose oxide has acidic properties (i.e.: Cr, Mn, V).

Oxoacids have a formula **HXO**, whereas hydroxides are indicated as **XOH**.

This convention allows to assign the properties of ternary compounds.

Nomenclature highlights the degree of oxidation:

suffix **-ous** = smaller oxidation number

suffix **-ic** = larger oxidation number

e.g.: sulphor**ous** acid (H_2SO_3 , o.n. S = **+4**)

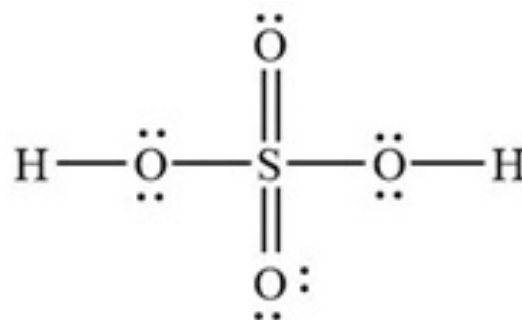
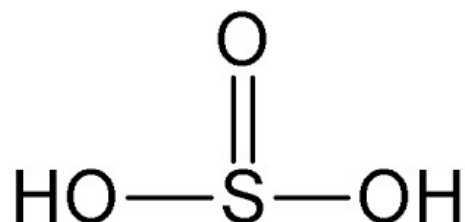
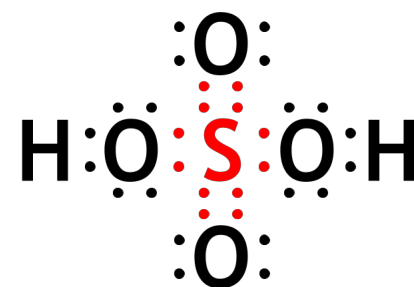
 sulphur**ic** acid (H_2SO_4 , n.o. S = **+6**)

Nitrous acid (HNO_2 , n.o. N = +3)

Nitric acid (HNO_3 , n.o. N = +5)

Sulphurous acid (H_2SO_3 , n.o. S = +4)

Sulphuric acid (H_2SO_4 , n.o. S = +6)



If suffixes do not allow to name all oxoacids prefixes are used:
hypo- and **per-**

hypochlorous acid (HClO , n.o. +1)

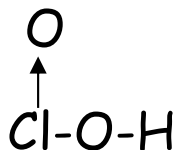
chlorous acid (HClO_2 , n.o. +3)

chloric acid (HClO_3 , n.o. +5)

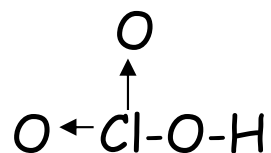
perchloric acid (HClO_4 , n.o. +7)



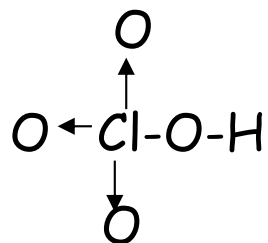
Hypochlorous acid



Chlorous acid



Chloric acid

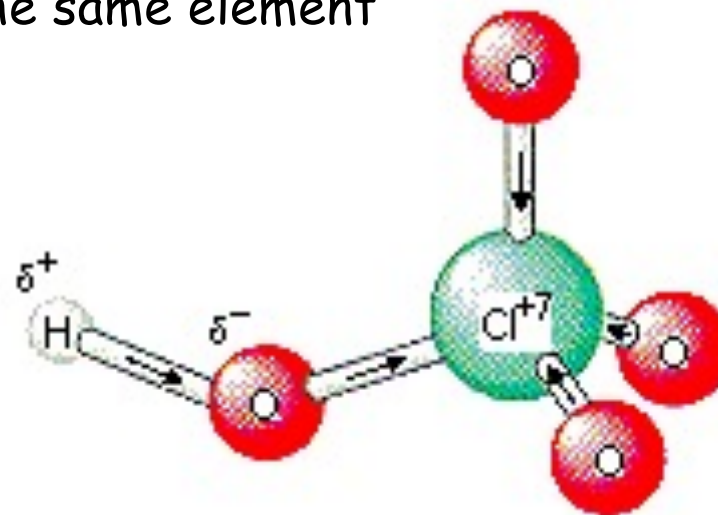


Perchloric acid

Oxygen atoms are added by means of coordination bond.

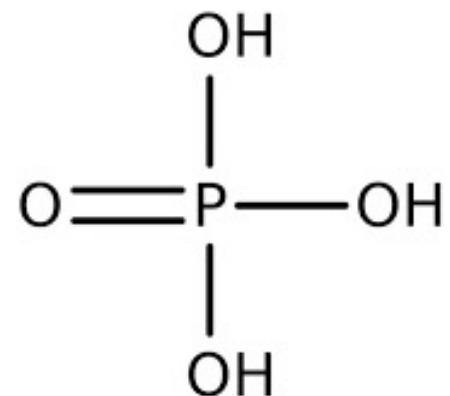
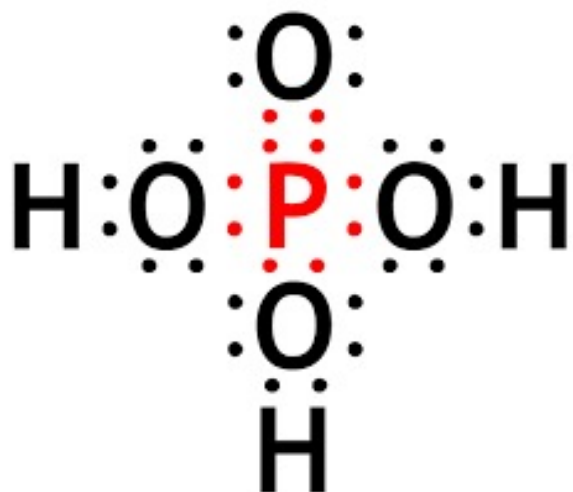
The maximum number of oxygen atoms depends the number of pairs available and the radius of the atom.

Behaviour of different oxoacids of the same element



Polarization of the O-H bond depends on the attraction of electron exerted by the O atom.

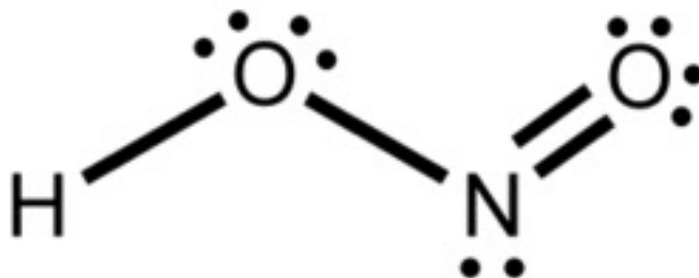
hypochlor**ous** acid (HClO , n.o. +1) Very weak
chlor**ous** acid (HClO_2 , n.o. +3) weak
chlor**ic** acid (HClO_3 , n.o. +5) strong
perchlor**ic** acid (HClO_4 , n.o. +7) very strong



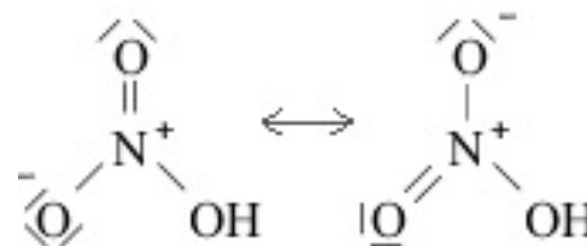
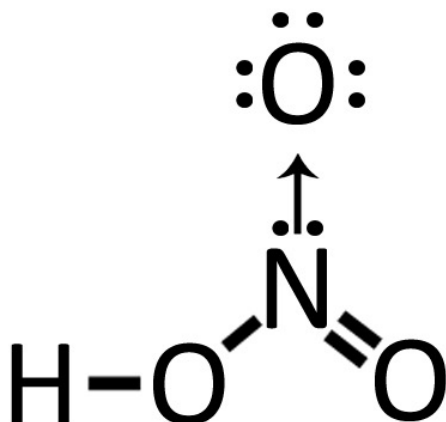
PHOSPHORIC ACID H_3PO_4

NITROUS AND NITRIC ACIDS

Nitrous acid

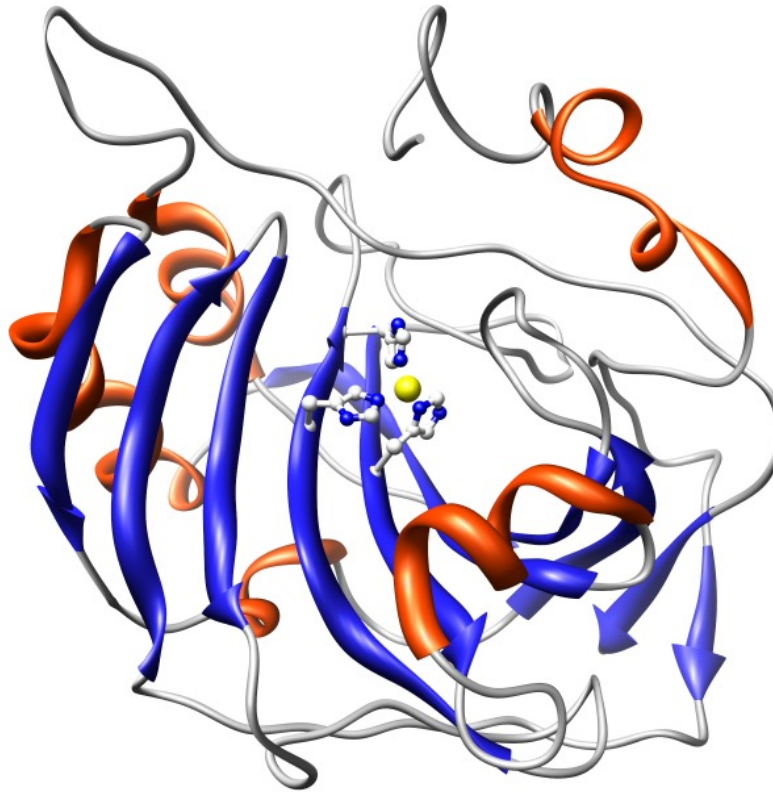


Nitric acid



Carbonic anhydrase plays a role in CO₂ transport in blood.

Venous blood: $\text{CO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{CO}_3 \rightarrow \text{HCO}_3^- + \text{H}^+$



Arterial blood: $\text{HCO}_3^- + \text{H}^+ \rightarrow \text{H}_2\text{CO}_3 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$

Common oxoacids (oxygen o.n. -2)

Oxoacid	Formula	Anion	Anion Formula
acetic acid	CH_3COOH	acetate	CH_3COO^-
carbonic acid	H_2CO_3	carbonate	CO_3^{2-}
chloric acid	HClO_3	chlorate	ClO_3^-
chlorous acid	HClO_2	chlorite	ClO_2^-
hypochlorous acid	HClO	hypochlorite	ClO^-
iodic acid	HIO_3	iodate	IO_3^-
nitric acid	HNO_3	nitrate	NO_3^-
nitrous acid	HNO_2	nitrite	NO_2^-
perchloric acid	HClO_4	perchlorate	ClO_4^-
phosphoric acid	H_3PO_4	phosphate	PO_4^{3-}
phosphorous acid	H_3PO_3	phosphite	PO_3^{3-}
sulfuric acid	H_2SO_4	sulfate	SO_4^{2-}
sulfurous acid	H_2SO_3	sulfite	SO_3^{2-}

Hydrides

Binary compounds of hydrogen. The ratio with elements from I to VII group are fixed: 1, 2, 3, 4, 3, 2, 1 (LiH, CaH₂, AlH₃, CH₄, NH₃, H₂O, HF)

Metal hydrides: with elements from groups IA and IIA. Hydrogen has o.n. -1. With the exception of LiH e BeH₂, they are ionic compounds. Hydrogen takes the form of **hydride ion** (H⁻). $H^- + H^+ \rightarrow H_2$; $\Delta H = -1676 \text{ kJ/mol}$

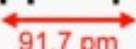
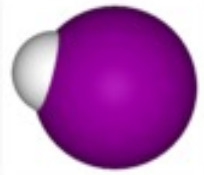
Covalent hydrides: with elements of groups from IV on (CH₄, methane; SiH₄, silane; NH₃, ammonia; PH₃, phosphane (phosphine)).

Hydracids : binary compounds of H with elements from groups VI and VII. Nomenclature IUPAC Hydrogen + name of the element + suffix **-ide**) (prefix **hydro-** + name of the element + acid):
Es. HF, **hydrofluoric acid**; HCl, **hydrochloric acid**; HBr, **hydrobromic acid**; HI, **hydroiodic acid**.
Hydrogen sulphide H₂S

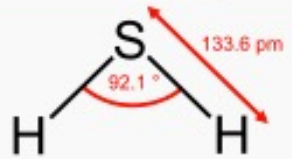
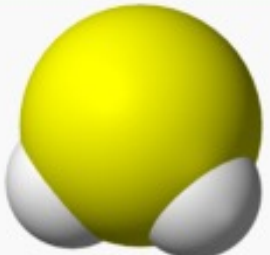
Cations from hydrides take the suffix **-onium**.

E.g.: PH₄⁺, **phosphonium**; NH₄⁺, **ammonium**; H₃O⁺, **oxonium** (or hydronium).

hydrogen halides

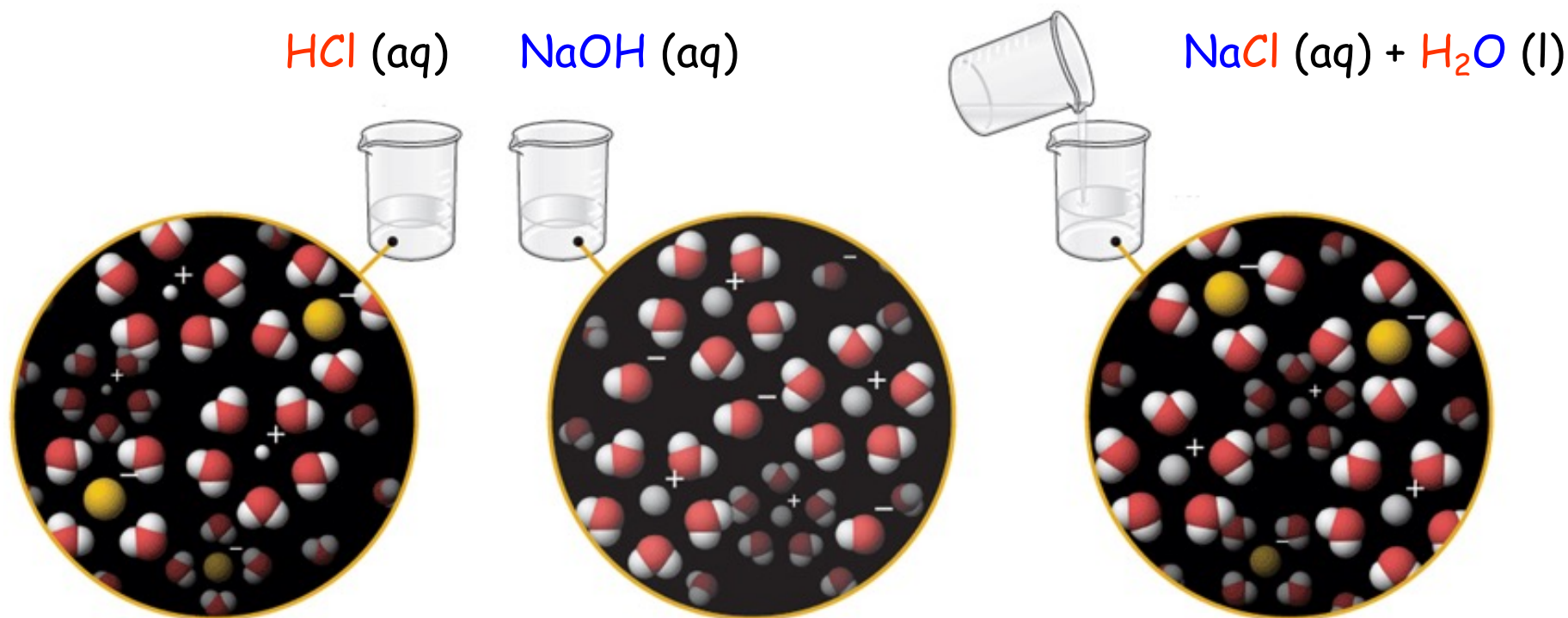
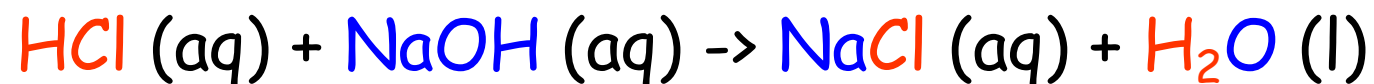
compound	formula	structure	model	$d(\text{H-X}) / \text{pm}$ (gas phase)	μ / D
hydrogen fluoride	HF	$\text{H}-\text{F}$ 		91.7	1.86
hydrogen chloride	HCl	$\text{H}-\text{Cl}$ 		127.4	1.11
hydrogen bromide	HBr	$\text{H}-\text{Br}$ 		141.4	0.788
hydrogen iodide	HI	$\text{H}-\text{I}$ 		160.9	0.382

Hydrogen oxide (H_2O): water.

Hydrogen sulfide	
	
Systematic name [hide] Hydrogen sulfide ^[1]	

Salts

Ionic compounds arising from the reaction of a **base** with an **acid** (neutralization of acids with bases).



Nomenclature of salts

The **metallic component** (positive) is written first, the **nonmetallic component** (acidic radical, negative) follows.

The name of the salt is formed with: cation name + acidic radical.

type	Acid suffix	Salt suffix	acid	salt
oxoacid	-ous	-ite	HNO ₂ Nitrous acid	NaNO ₂ Sodium nitrite
oxoacid	-ic	-ate	H ₂ SO ₄ Sulphuric acid	CaSO ₄ Calcium sulphate
hydracid	-ide	-ide	HCl Hydrogen chloride	KCl Potassium chloride

common salts

acid	base	salt	IUPAC name	Traditional name
HCl	NaOH	NaCl	Sodium chloride	Sodium chloride
HF	Ca(OH) ₂	CaF ₂	Calcium difluoride	Calcium fluoride
HBr	Al(OH) ₃	AlBr ₃	Aluminum tribromide	Aluminum bromide
H ₂ S	Fe(OH) ₂	FeS	Iron (II) sulphide	Ferrous sulphide
H ₂ S	Fe(OH) ₃	Fe ₂ S ₃	Iron (III) trisulphide	Ferric sulphide
H ₂ S	KOH	K ₂ S	Dipotassium sulphide	Potassium sulphide
H ₂ CO ₃	NaOH	Na ₂ CO ₃	Disodium carbonate	Sodium carbonate
H ₂ SO ₄	KOH	K ₂ SO ₄	Dipotassium sulphate	Potassium sulphate
H ₂ SO ₄	Fe(OH) ₂	FeSO ₄	Iron (II) sulphate	Ferrous sulphate
H ₂ SO ₄	Fe(OH) ₃	Fe ₂ (SO ₄) ₃	Iron (III) trisulphate	Ferric sulphate
H ₃ PO ₄	Ca(OH) ₂	Ca ₂ (PO ₄) ₃	Dicalcium triphosphate	Calcium phosphate
HMnO ₄	KOH	KMnO ₄	Potassium manganate(IV)	Potassium permanganate

Ionic compounds

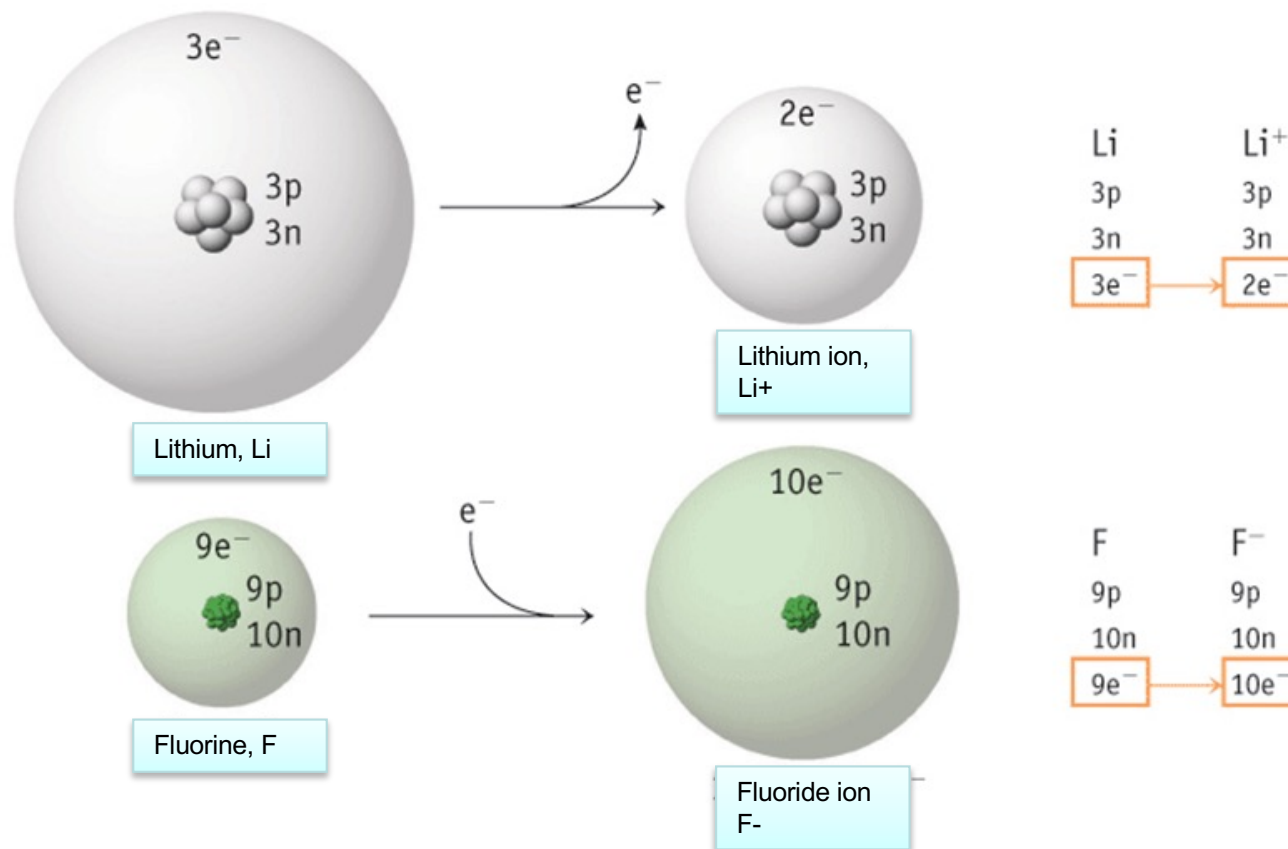
Ionic compounds consist of ions, atoms or groups of atoms that have positive or negative charges.

Common name	name	formula	Ions
salt	Sodium chloride	NaCl	Na^+ , Cl^-
lime	Calcium oxide	CaO	Ca^{2+} , O^{2-}
limestone	Calcium carbonate	CaCO_3	Ca^{2+} , CO_3^{2-}
fluorite	Calcium fluoride	CaF_2	Ca^{2+} , F^-
gypsum	Calcium sulphate dihydrate	$\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$	Ca^{2+} , SO_4^{2-}
hematite	Iron oxide (III)	Fe_2O_3	Fe^{3+} , O^{2-}
orpiment	Arsenic sulfide	As_2S_3	As^{3+} , S^{2-}



The atoms of many elements **can give away or acquire electrons** during a chemical reaction. To be able to **predict the outcome of a chemical reaction**, one needs to know if an element acquires or gives up electrons, and if so, how many.

When an atom **gives away one or more electrons** (transferred to another atom in a reaction), a positively charged ion called **cation** is formed. When an atom **acquires an electron**, a negatively charged ion called **anion** is formed.



Can we predict if an atom of an element will preferentially form a cation or an anion?

- **Metals** generally give up their electrons giving rise to **cations**.
- **Non-metals** frequently acquire electrons giving rise to **anions**.

1A																	8A
1 H	2A											3A	4A	5A	6A	7A	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 –	113 –	114 –	115 –	116 –		

metals

Transition metals

metalloids

Non-metals
35

Mono-atomic ions

From elements which have lost or acquired electrons.

1A-3A form positive ions with charge equal to the number of the group.

There are no rules for transition metals.

Non-metals often form ions with negative charge equal to: 8- group number.

1A	2A											3A	4A	5A	6A	7A	8A
H ⁺														N ³⁻	O ²⁻	F ⁻	
Li ⁺												Al ³⁺		p ³⁻	S ²⁻	Cl ⁻	
Na ⁺	Mg ²⁺	3B	4B	5B	6B	7B	8B		1B	2B					Se ²⁻	Br ⁻	
K ⁺	Ca ²⁺		Ti ⁴⁺		Cr ²⁺	Mn ²⁺	Fe ²⁺	Co ²⁺	Ni ²⁺	Cu ⁺					Te ²⁻	I ⁻	
Rb ⁺	Sr ²⁺				Cr ³⁺		Fe ³⁺	Co ³⁺		Cu ²⁺	Zn ²⁺						
Cs ⁺	Ba ²⁺									Ag ⁺	Cd ²⁺		Sn ²⁺				
											Hg ₂ ²⁺		Pb ²⁺	Bi ³⁺			
											Hg ²⁺						

group	element	electrons	cation
metals			
1A	Na (11 protons, 11 electrons)	-1	Na ⁺ (11 protons, 10 electrons)
2A	Ca (20 protons, 20 electrons)	-2	Ca ²⁺ (20 protons, 18 electrons)
3A	Al (13 protons, 13 electrons)	-3	Al ³⁺ (13 protons, 10 electrons)
Transition metals			
7B	Mn (25 protons, 25 electrons)	-2	Mn ²⁺ (25 protons, 23 electrons)
8B	Fe (26 protons, 26 electrons)	-2	Fe ²⁺ (26 protons, 24 electrons)
8B	Fe (26 protons, 26 electrons)	-3	Fe ³⁺ (26 protons, 23 electrons)
Non-metals			
5A	N (7 protons, 7 electrons)	+3	N ³⁻ (7 protons, 10 electrons)
6A	S (16 protons, 16 electrons)	+2	S ²⁻ (16 protons, 18 electrons)
7A	Br (35 protons, 35 electrons)	+1	Br ⁻ (35 protons, 36 electrons)

Notation for positive ions (cations).

With few exceptions (such as NH_4^+ , ammonium) positive ions arise from metals. They are named according to these rules.

1) For a positive **mono-atomic** ion, use the name of the **metal** followed by the word "**cation**". E.g.. Ag^+ is called "Silver cation".

2) Cations that can have on **more than one positive charge** are labeled with Roman numerals in parentheses. E.g.: Co^{1+} is Copper (I) and Cu^{2+} is Copper(II) ("copper one cation" and "copper two cation").

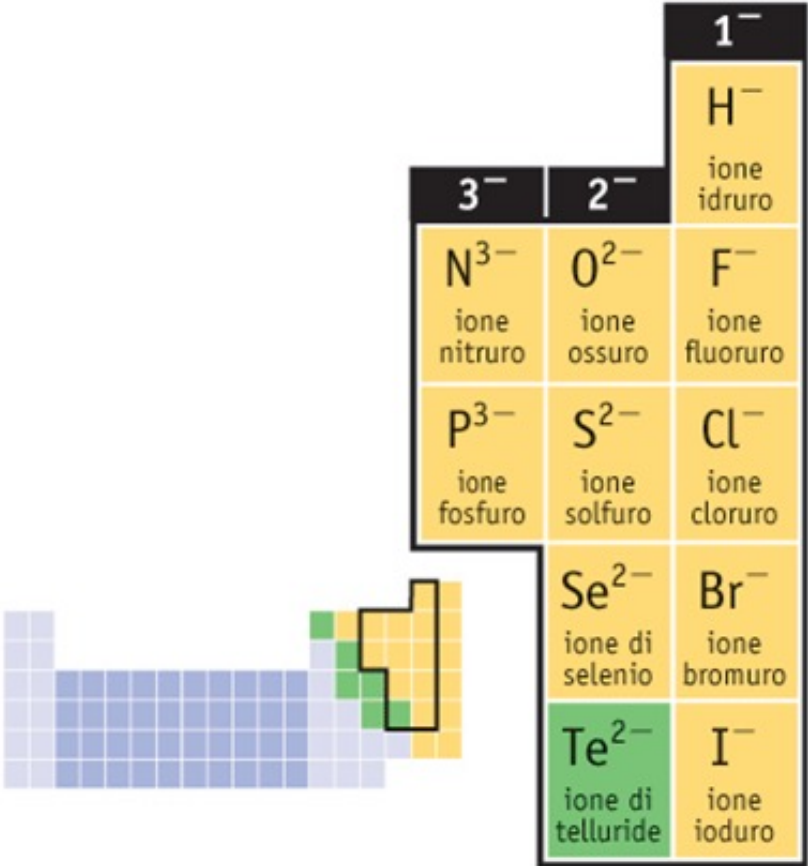
1) An older, notation is to append **-ous** or **-ic** to the root of the Latin name to name ions with a lesser or greater charge: Cu^{2+} = **Cupric** and Cu^{1+} = **Cuprous**.

Notation for negative ions (anions).

Two types of anions: monoatomic and polyatomic

The name of a **monoatomic** negative ion is formed by adding the **-ide** ending to the name of the non-metal element. E.g. Cl^- is called **chloride** anion.

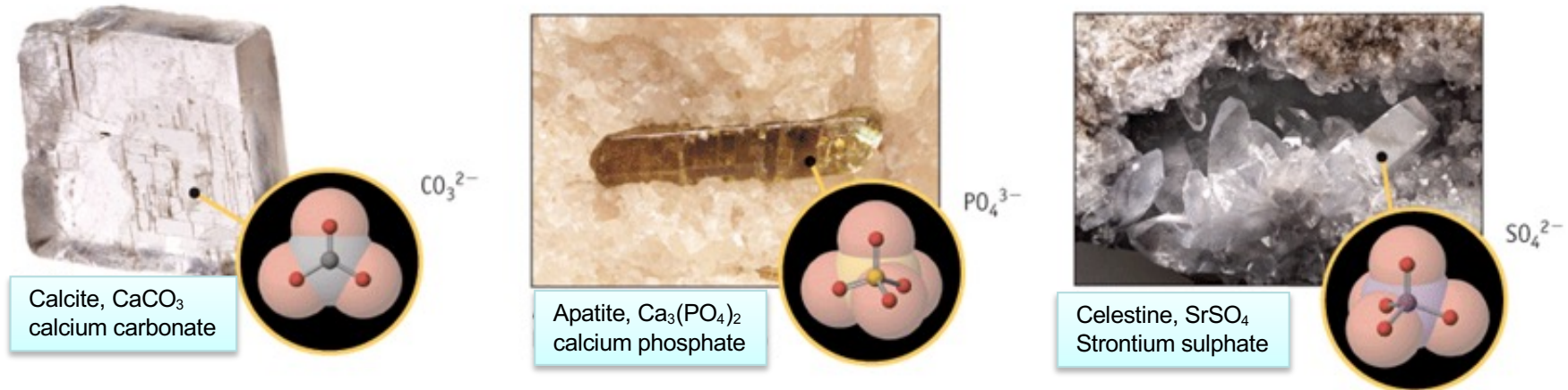
H^- Hydride
 F^- Fluoride
 O^{2-} Oxide
 S^{2-} Sulfide
 N^{3-} Nitride
 P^{3-} Phosphide



1⁻		
H^- ione idruro		
3⁻	2⁻	
N^{3-} ione nitruro	O^{2-} ione ossuro	F^- ione fluoruro
P^{3-} ione fosfuro	S^{2-} ione solfuro	Cl^- ione cloruro
	Se^{2-} ione di selenio	Br^- ione bromuro
	Te^{2-} ione di telluride	I^- ione ioduro

Molecular ions (polyatomic)

Made up by more than one atom, electrically charged.



E.g.: the carbonate ion, CO_3^{2-} , is formed by 1 atom of C and 3 atoms of O. It has 2 negative charges since it contains two extra e^- .

Ionic compounds have net charge=0

Ruby is made up by ions Al^{3+} and oxide ions O^{2-} . To achieve null charge 2 Al^{3+} ions [total charge = $2 \times (3+) = 6+$] are combined with 3 O^{2-} ions [total charge = $3 \times (2-) = 6+$] to yield Al_2O_3 .



Polyatomic ions that contain oxygen are called **oxyanions**.

- Oxyanions are named with **-ite** or **-ate**, for a lesser or greater quantity of oxygen.

NO_3^- is called nit**ate** and NO_2^- is called nit**ite**

SO_4^{2-} is called sulph**ate** e SO_3^{2-} is called sulph**ite**

- If four oxyanions are possible, the prefixes **hypo-** and **per-** are used:

ClO_4^- is called **perchlorate** e ClO_3^- is called chlor**ate**

ClO_2^- is called chlor**ite** e ClO^- is called **hypochlorite**

- The modern systematic specifically names the hydrogen atom.

HPO_4^{2-} is **hydrogen** phosphate e $\text{H}_2\text{PO}_4^{1-}$ is **dihydrogen** phosphate

HCO_3^- is called **hydrogen** carbonate (bicarbonate) e HSO_4^{2-} is **hydrogen** sulphate.

Formula	Name	Formula	Name
Group 4A		Group 7A	
CN ⁻	Cyanide ion	ClO ⁻	Hypochlorite ion
CH ₃ CO ₂ ⁻	Acetate ion	ClO ₂ ⁻	Chlorite ion
CO ₃ ²⁻	Carbonate ion	ClO ₃ ⁻	Chlorate ion
HCO ₃ ⁻	Hydrogen carbonate (bicarbonate)	ClO ₄ ⁻	Perchlorate ion
Group 5A		Transition metals	
NO ₂ ⁻	Nitrite ion	CrO ₄ ⁻	Chromate ion
NO ₃ ⁻	Nitrate ion	Cr ₂ O ₇ ²⁻	Dichromate ion
PO ₄ ³⁻	Phosphate ion	MnO ₄ ⁻	Permanganate ion
HPO ₄ ²⁻	Hydrogen phosphate ion		
H ₂ PO ₄ ⁻	Di-hydrogen phosphate		
Group 6A			
OH ⁻	Hydroxide ion		
SO ₃ ²⁻	Sulphite ion		
SO ₄ ²⁻	Sulphate ion		
HSO ₄ ⁻	Hydrogen sulphate ion		

Inorganic compounds

ACIDS

1. Hydrogen halides

HF	hydrogen fluoride (hydrofluoric acid)	strong	caustic
HCl	hydrogen chloride (hydrochloric acid)	strong	digestive fluid
HBr	hydrogen bromide (hydrobromic acid)	strong	caustic
HI	hydrogen iodide (hydroiodic acid)	strong	caustic

2. Oxyacids

HClO	hypochlorous acid	weak	oxidant
HClO ₂	chlorous acid	weak	oxidant
HClO ₃	chloric acid	strong	oxidant
HClO ₄	perchloric acid	strong	oxidant/caustic
HNO ₂	Nitrous acid	weak	atmosphere/ozone depletion
HNO ₃	nitric acid	strong	caustic
H ₂ SO ₄	sulfuric acid	strong	caustic
H ₂ SO ₃	sulphorous acid	strong	caustic
H ₃ PO ₄	phosphoric acid	weak	buffer in blood
H ₃ BO ₃	boric acid	weak	antiseptic
H ₂ CO ₃	carbonic acid	weak	buffer in blood
HMnO ₄	Permanganic acid		oxidizing salts
HCN	hydrogen cyanide (cyanidric acid)	weak	poison

Inorganic compounds

BASES (hydroxides)

NaOH	sodium hydroxide	strong	caustic
KOH	potassium hydroxide	strong	caustic
Ca(OH) ₂	calcium hydroxide	strong	caustic
Mg(OH) ₂	magnesium (di)hydroxide	strong	caustic
Al(OH) ₃	aluminum (tri)hydroxide	strong	caustic
NH ₃ .H ₂ O	ammonia	weak	surface cleaning
NH ₄ OH	ammonium hydroxide		

OXIDES

CO	carbon monoxide	lethal gas
Na ₂ O	(di)sodium oxide	yields hydroxide in water
K ₂ O	(di)sodium oxide	yields hydroxide in water
CaO	calcium oxide	yields hydroxide in water
MgO	magnesium oxide	yields hydroxide in water
Al ₂ O ₃	(di)aluminum (tri)oxide	yields hydroxide in water

OXIDES (anhydrides that in water yield oxyacids)

CO ₂	carbon dioxide	present in air, cell resp. product
NO	nitrogen oxide	messenger, vasodilation, toxic
NO ₂	nitrogen dioxide	$2 \text{NO}_2 + \text{H}_2\text{O} \rightarrow \text{HNO}_2 + \text{HNO}_3$
SO ₂	sulfur dioxide	Exhaustion gas (+ H ₂ O acid rain) $\text{SO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_3$
SO ₃	sulfur trioxide	$\text{SO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_4$

Inorganic compounds

SALT

Non-hydrolizing salts (neutral)

NaCl	sodium chloride	body fluids, cells
KCl	potassium chloride	ubiquitario
KBr	potassium bromide	sedative
KI	potassium iodide	supply for thyroid
CaCl ₂	calcium chloride	ion supply
MgSO ₄	magnesium sulfate	low solubility/ laxative
KNO ₃	potassium nitrate	fertilizer/agriculture
NaNO ₃	sodium nitrate	fertilizer/agriculture
CuSO ₄	copper sulfate	antimicrobial/agriculture
AgNO ₃	silver nitrate	local antiseptic
KMnO ₄	potassium permanganate	antiseptic/oxidizing
HgCl ₂	mercury choride	toxic
BaSO ₄	barium sulfate	low solubility/ x-ray diagnostics

SALT (Acidic hydrolysis)

NH ₄ Cl	ammonium chloride	urine
CH ₃ COO(NH ₄)	ammonium acetate	lab reagent
(NH ₄) ₂ SO ₄	ammonium sulphate	lab reagent

SALT*(basic hydrolysis)*

CH_3COONa	sodium acetate	in buffers
CH_3COOK	potassium acetate	in buffers
CaCO_3	calcium carbonate	antacid
$\text{Ca}_3(\text{PO}_4)_2$	calcium phosphate	bone matrix
Na_3PO_4	sodium phosphate	industry
Na_2HPO_4	disodium hydrogen phosphate	buffer in plasma
NaH_2PO_4	mono-sodium hydrogen phosphate	"
NaHCO_3	sodium hydrogencarbonate	<i>(bicarbonate) blood</i>

Coordination complex

Compound in which an atom or ion is bound to other chemical species with a number larger than its oxidation number.

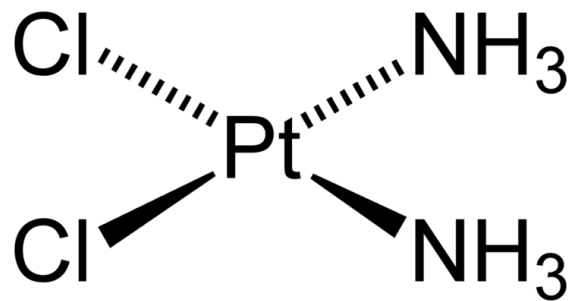
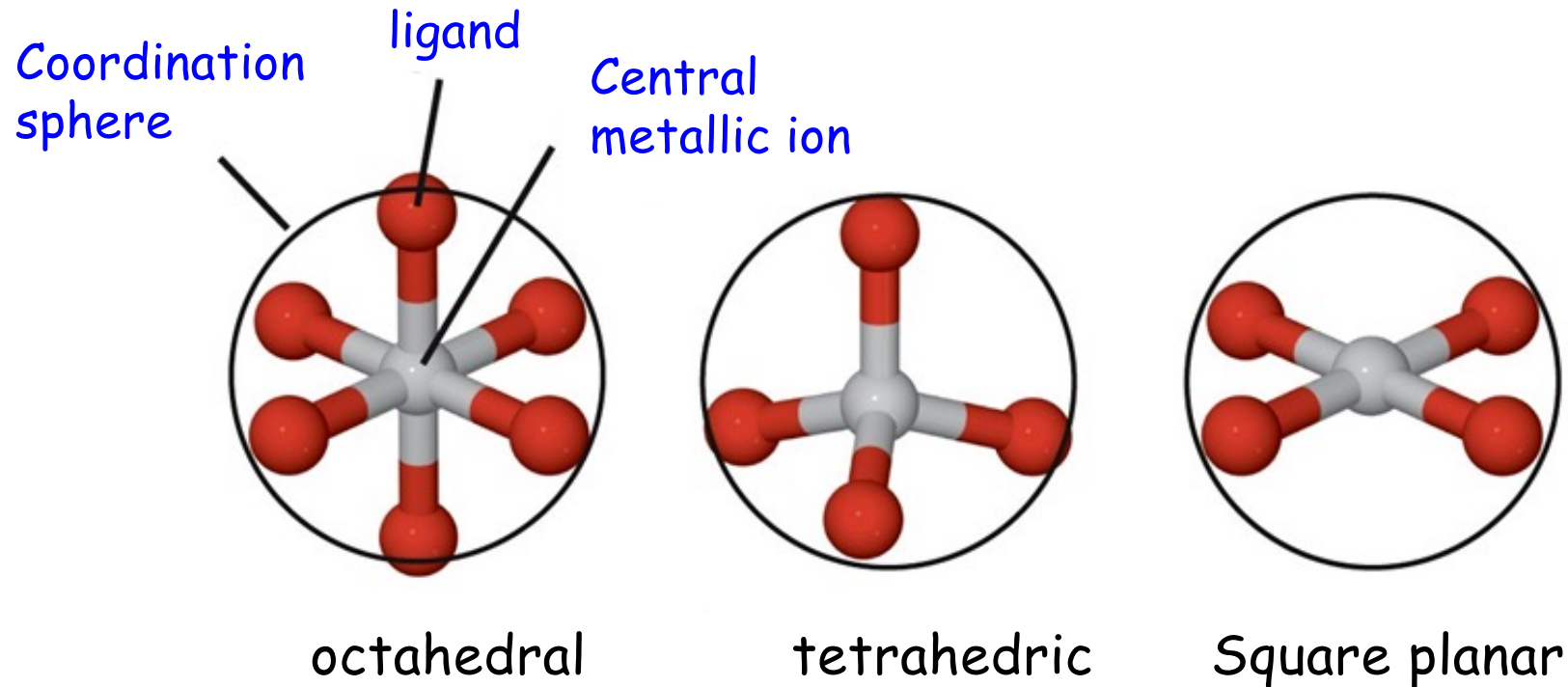
Typically the central atom is a **transition metal cation**, the ion or molecules that surround it are called **ligands**.

The set of ligand forms the **coordination sphere** of the complex and the number of ligands is called the **coordination number** (ranging from 1 to 16).

The bond in the coordination complexes arises from empty d orbital of the central atoms and lone pairs from s and p orbitals of the ligands.

The resulting bonds have transition energies that fall in the **visible** region of the spectrum, therefore they have characteristic colours (heme and chlorophyll).

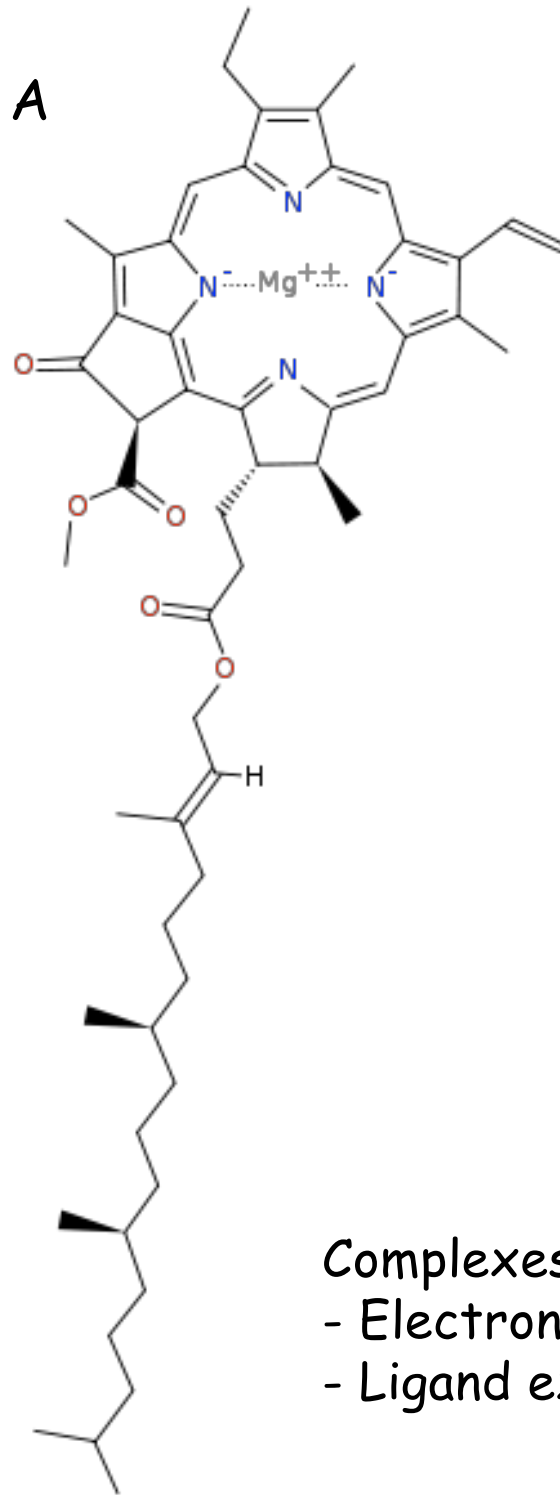
E.g. a metal in aqueous solution (coordinated by water or other ligands).



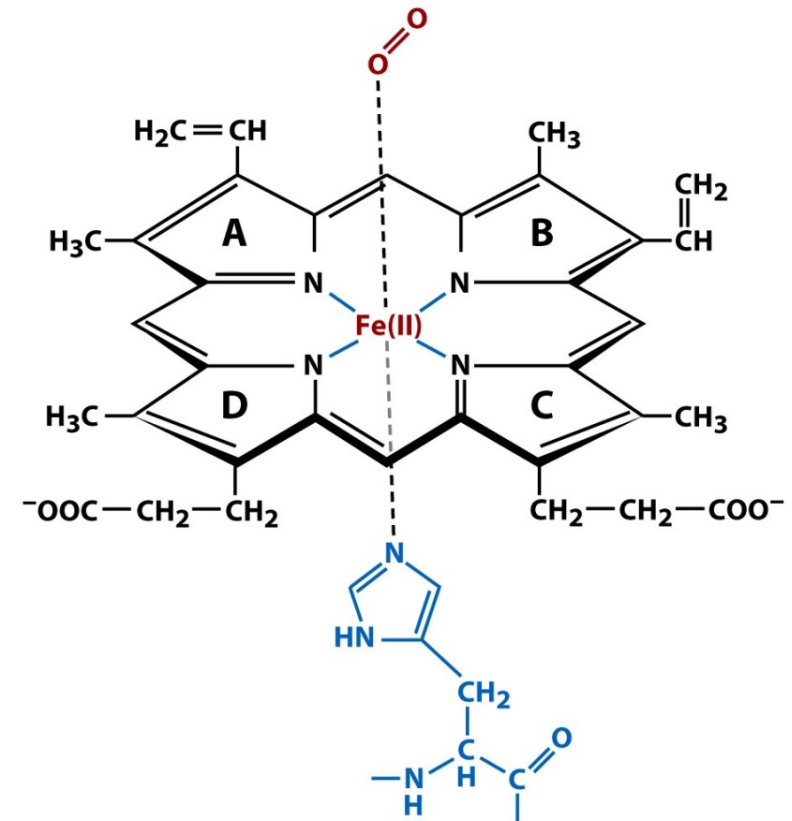
Cisplatinum(I), a drug used in chemotherapy.

It reacts *in vivo*, binding to and causing crosslinking of DNA, triggering apoptosis (programmed cell death).

Chlorophyll A



Heme in hemoglobin
and myoglobin



Complexes show a variety of possible reactivities:

- Electron transfers
- Ligand exchange

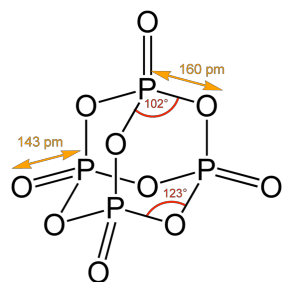
Meta-, ortho- e pyro- acids

This (obsolete) nomenclature is still used to distinguish the degree of hydration of an oxoacid.

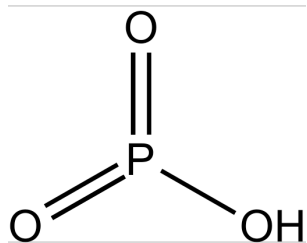
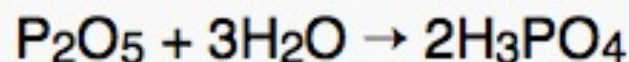
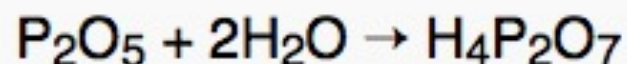
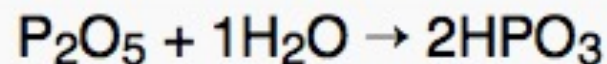
Ortho- the acid containing the largest number of water molecules.

Pyro- the acid containing one less water molecule

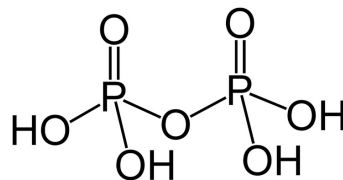
Meta- the acid containing the fewer number of water molecules



Phosphorous pentoxide

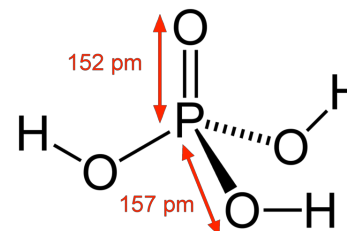


metaphosphoric acid
(HPO_3)



Pyrophosphoric acid
($\text{H}_4\text{P}_2\text{O}_7$)

Phosphoric acid



Orthophosphoric acid
(H_3PO_4)

Inorganic compounds nomenclature

Currently known inorganic compounds are about 6 million and their number increases by about 6000 a week. Such a number of substances need to be organized according to clear, simple and universally shared rules.

The purpose of nomenclature is to provide rules for identifying a compound, giving it a **unique** name and a **formula**, using as few words as possible

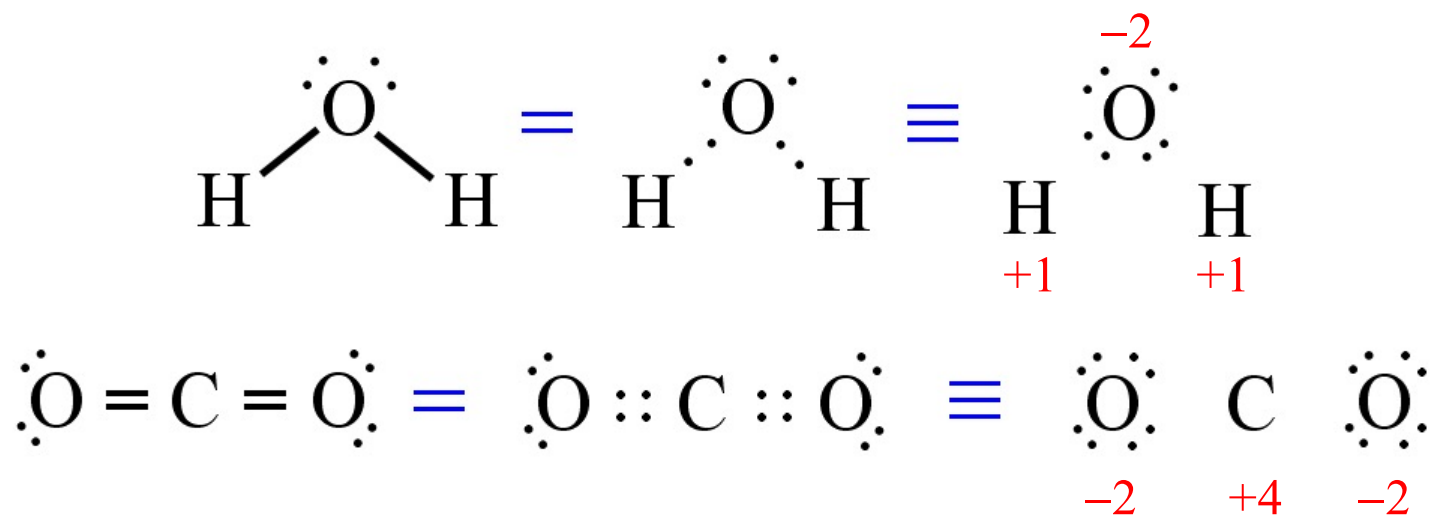
Rules for nomenclature are published by IUPAC (International Union of Pure and Applied Chemistry). <http://www.iupac.org/> "Traditional" nomenclature is sometimes still in use.



Oxidation number

O.N. is the charge that would be acquired by an element in a compound if bonding electrons were on the more electronegative atom.

The O.N. is deduced from its external electronic configuration in its fundamental state.

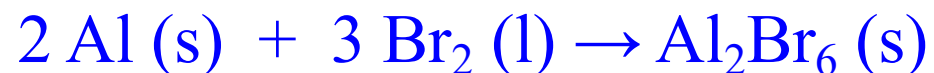


The oxidation number is not an actual charge, but it is only formally attributed to the atom.

The oxidation numbers allow to **distribute electrons** between the atoms of a molecule. Since the distribution of electrons changes during a **redox** reaction, this method is used to identify a redox reaction, the **reducing and oxidizing** agents and to balance the chemical equation.



Aluminum and boron react to form aluminum bromide.



n.o. = 0

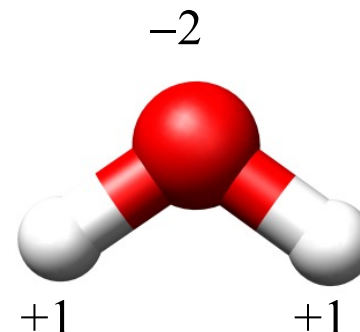
n.o. = 0

n.o. Al = +3 e n.o. Br = -1

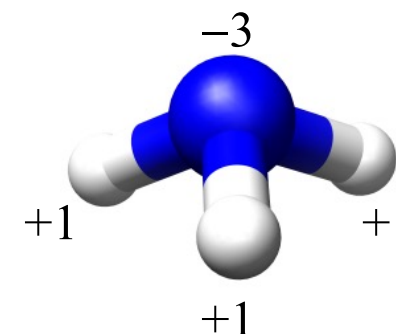
To calculate the oxidation number in a compound:

- 1. **Identify** the most electronegative atom
- 2. **Assign** the electrons involved in bonds
- 3. Evaluate the gain and loss of electrons

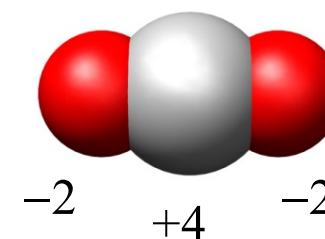
H ₂ O	Z	electronegativity	Electron configuration	Oxidation number
H	1	2.2	1s ¹	+1 (1s ⁰)
O	8	3.5	1s ² 2s ² 2p ⁴	-2 (1s ² 2s ² 2p ⁶)



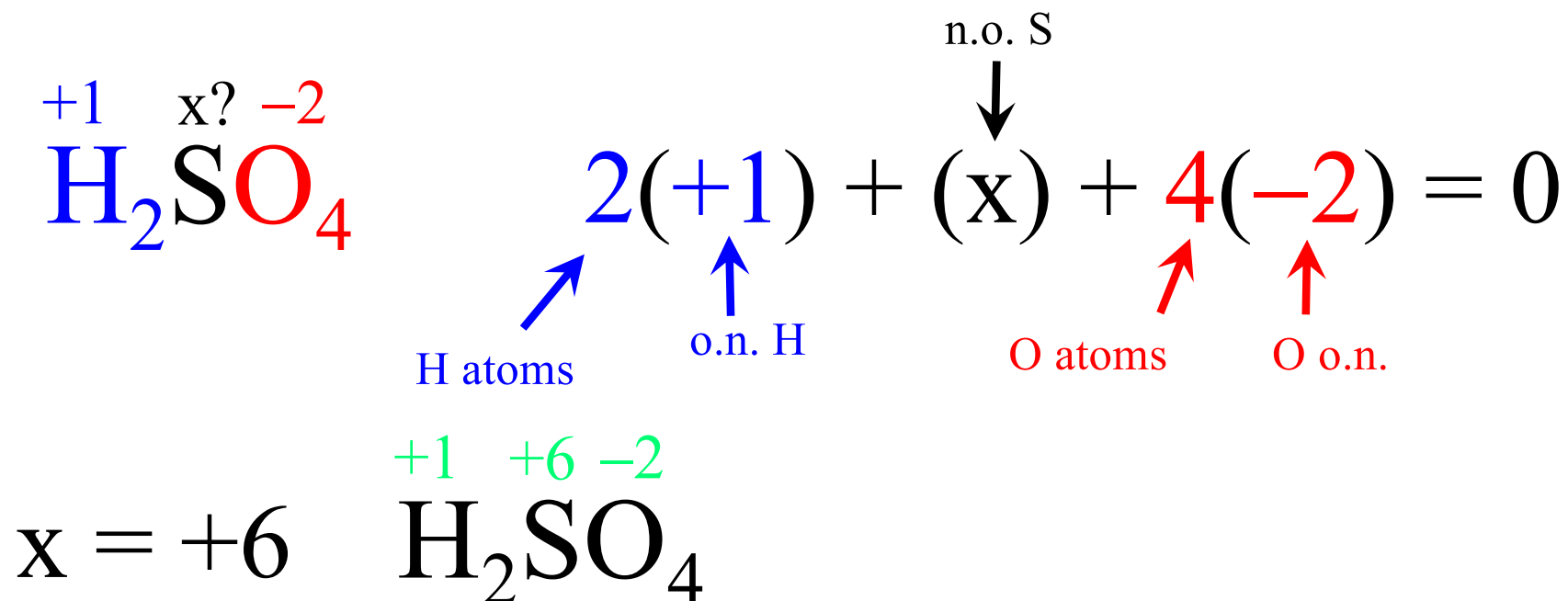
NH ₃	Z	electronegativity	Electron configuration	Electron configuration
H	1	2.2	1s ¹	+1 (1s ⁰)
N	7	3.0	1s ² 2s ² 2p ³	-3 (1s ² 2s ² 2p ⁶)



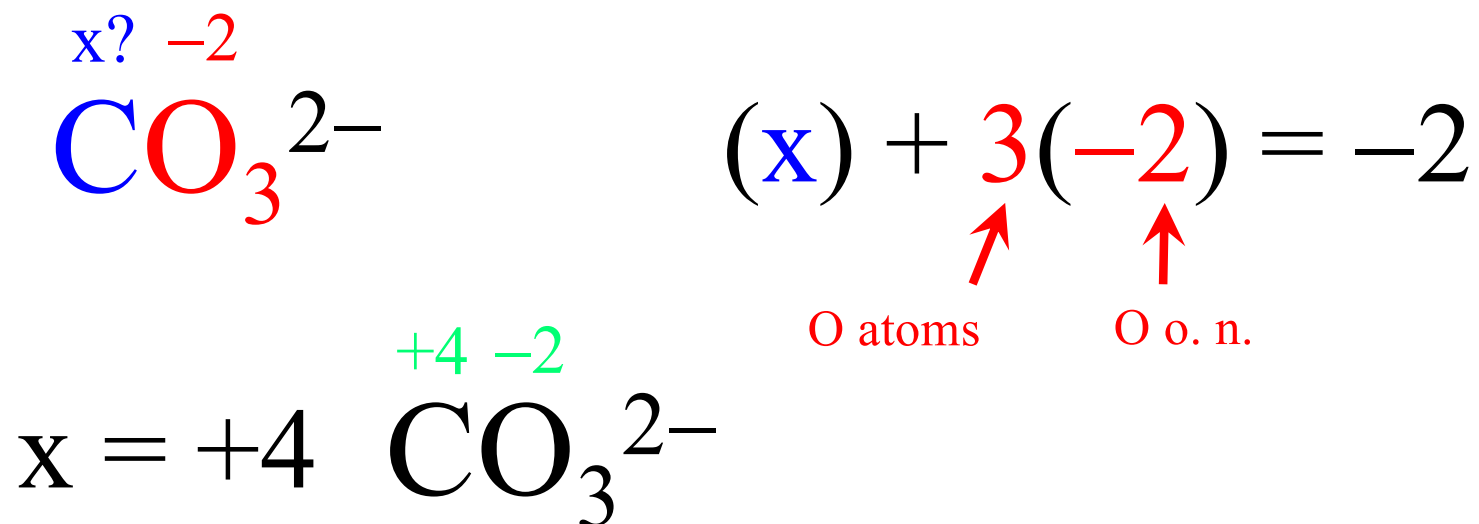
CO ₂	Z	electronegativity	Electron configuration	Electron configuration
C	6	2.5	1s ² 2s ² 2p ²	+4 (1s ²)
O	8	3.5	1s ² 2s ² 2p ⁴	-2 (1s ² 2s ² 2p ⁶)



1. What is the o.n. of sulphur in sulphuric acid?



2. What is the o.n. in the anion carbonate?



Determine the oxidation number of:

- a) Aluminum in aluminum oxide, Al_2O_3
- b) Phosphorus in phosphoric acid H_3PO_4
- c) Sulphur in sulphate SO_4^{2-}

a) Al_2O_3 bears no charge, $\text{O} = -2$, $\text{Al} = +3$ (gruppo 3A):

$$\text{Net charge } \text{Al}_2\text{O}_3 = 0$$

$$2 (+3) + 3 (-2) = 0$$

b) H_3PO_4 bears no charge. If O has o.n. = -2 , and each H atom has o.n. = $+1$, il P must have o.n. = $+5$

$$\text{Net charge } \text{H}_3\text{PO}_4 = 0$$

$$3 (+1) + (+5) + 4 (-2) = 0$$

c) Sulphate ion has a 2- charge SO_4^{2-} $\text{O} = -2$, therefore $\text{S} = +6$.

$$\text{Net charge } \text{SO}_4^{2-} = 2-$$

$$\text{o.n. S} + \text{n.o. O} = -2$$

$$(+6) + 4 (-2) = -2$$