Naming Inorganic Compounds
Ionic Equations

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Atom

- Smallest particle of a pure element having its chemical properties
- Positively charged nucleus (protons, neutrons)
- Negatively charged electron shell:
  - Electron is wave/particle
  - Behavior of electron described by quantum mechanics (... wave function, quantum numbers)
  - Orbital: space area within the atom shell where occurrence of an electron or pair of electrons is more probable
- Structure of electron shell determines chemical properties
  - Valence electrons
  - The octet rule

Periodic table of elements

Group→ ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	55 Cs	56 Ba	*	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
7	87 Fr	88 Ra	**	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Fl	115 Uup	116 Lv	117 Uus	118 Uuo
		*	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	
		**	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	

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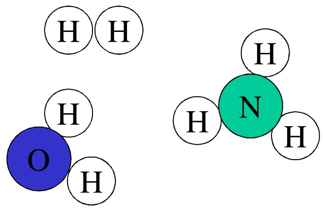
Compounds

- Molecular (covalent)
- Ionic

Molecular (covalent) compounds

- Consist of particles (molecules) made of several atoms connected with **covalent bonds**
- Examples of molecules:
  - Gases: diatomic
  - H_2O, NH_3 etc.



  - Covalently-bonded crystals: diamond
  - Macromolecules of proteins and nucleic acids

Diamond and graphite

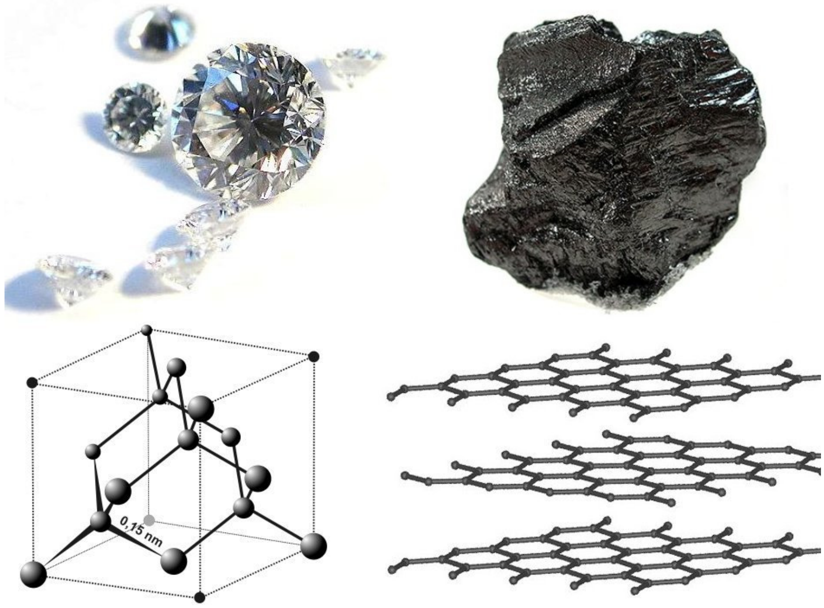
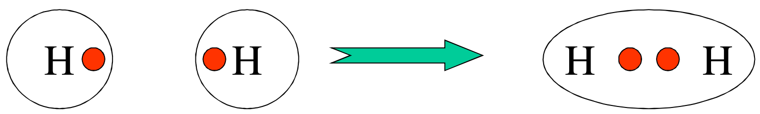


Fig. from Wikipedia

Covalent chemical bond

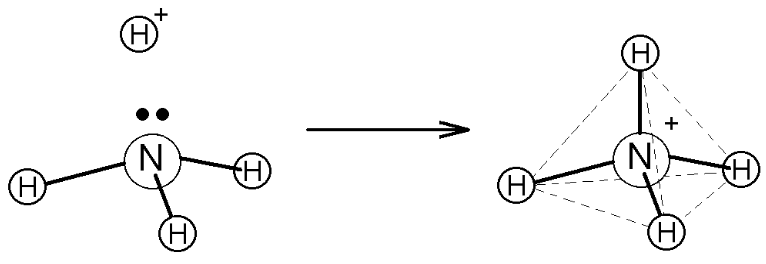
- Most often between two non-metals
- Chemical bond that results from two nuclei attracting the same pair(s) of electrons
- Based on electron sharing



- If more pairs shared … double or triple bonds
- Some atoms may also have nonbonding (lone) electron pairs

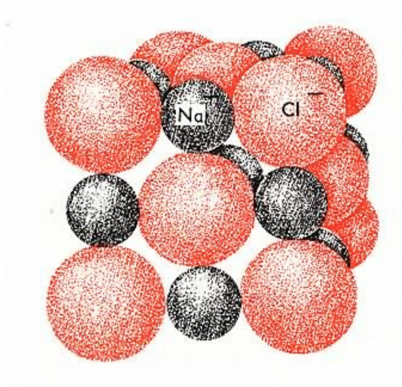
Coordination covalent bond

- (also dative, donor-acceptor bond)
- Both bonding electrons provided by one of the atoms (donor), whereas the other atoms provides an empty orbital (akceptor)



Ionic compounds

- Consist of charged particles (ions) held together by **ionic bonds**
- Metal + non-metal(s)
- Cations (+) and anions (-) combine in space to achieve electroneutrality

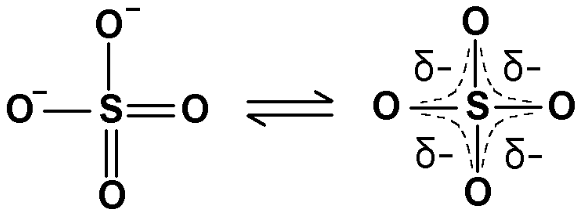


Ions

- Charged because number of electrons does not match number of protons
- Tendency to form ions depends on **electronegativity** of element
- Monoatomic: Na^{+}, Cl^{-}, H^{+}, Fe^{2+}
- Polyatomic: NO_3^{-} , SO_4^{2-}
- Complex: $[Fe(CN)_6]^{4-}$

Polyatomic ions of oxo-acids:

e.g. sulfate, SO_4^{2-} :



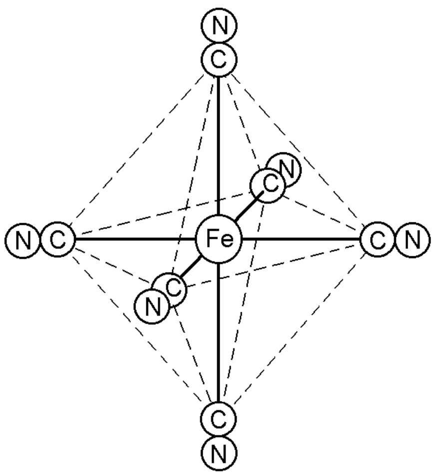
resonance stabilization of sulfate ion

..similar is nitrate NO_3^-, phosphate PO_4^{3-}, carbonate CO_3^{2-}, etc.

Coordination (complex) ions

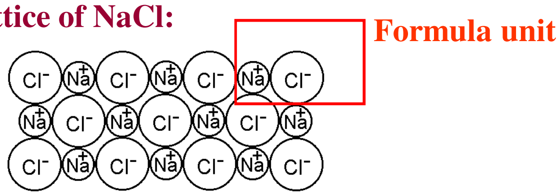
e.g. $[Fe(CN)_6]^{4-}$

- central atom of transition metal providing empty orbitals
- ligands providing free electron pairs
- Number of ligands (coordination number) is usually 4 or 6

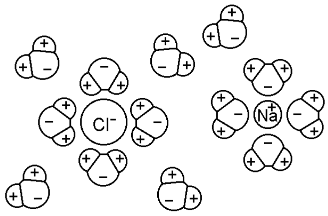


Ionic salts: no true molecule

- **Crystal lattice of NaCl:**



- **Dissolution of NaCl in water: electrolytic dissociation producing hydrated independent ions Na^+, Cl^-**



Polarity of chemical bond

Determined by difference in electronegativity of the two connected atoms:

< 0.4 nonpolar covalent bond

e.g. H-H, carbon-hydrogen

0.4 - 1.7 polar covalent bond

e.g. H-O-H, NH_3, carbon-oxygen, carbon-nitrogen

>1.7 ionic bond

e.g NaCl…

Gradual transition !

Oxidation number (formal valency)

- Oxidation number of element in compound equals its charge after giving all bonding electron pairs to the more electronegative atom
- Can be zero, positive or negative integer
- Basis for nomenclature of inorganic compounds
- Redox reactions: oxidation number increases in oxidation, decreases in reduction

Czech nomenclature of oxides:

Oxidation number	Suffix	General formula
I	-ný	X_2O
II	-natý	XO
III	-itý	X_2O_3
IV	-ičitý	XO_2
V	-ečný/-ičný	X_2O_5
VI	-ový	XO_3
VII	-istý	X_2O_7
VIII	-ičelý	XO_4

Rules for determination of oxidation numbers

- Free electroneutral atom, or atom in molecule of pure element: oxidation number = 0
- Oxidation number of a monoatomic ion equals its charge
- In heteroatomic compounds the bonding electrons are given to the more electronegative atom, practically:
  - H has nearly always oxidation number I (only in metallic hydrides -I)
  - O almost always -II (only in peroxides -I)
  - F always -I
  - Alkali metals (Na, K..) always I
  - Alkaline earth elements (Ca, Mg..) always II

Rules for determination of oxidation numbers

Examples:

CO_2 : C^{IV}, O^{-II}

H_2SO_4: H^{I}, S^{VI}, O^{-II}

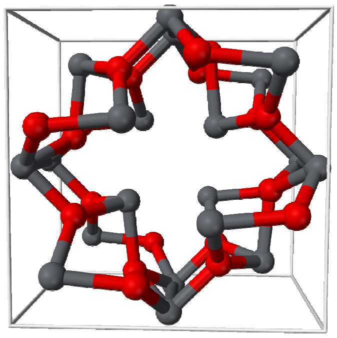
Sum of oxidation numbers of all atoms in electroneutral molecule is 0, in polyatomic ion equals the ion charge

e.g: CO_3^{2-}: C^{IV}, O^{-II} $\quad 1\times IV + 3 \times(-II) = -2$

What is oxidation number of lead in red lead (minium), Pb_3O_4?

II and IV

Red lead is chemically
2 [PbO] . [PbO_2]



Figures from Wikipedia

Naming Binary Ionic Compounds

- Compounds of one metallic element and one nonmetal (e.g. metallic oxides, hydroxides, halogenides, sulfides)
- Name of metal + stem of nonmetal + **-ide**
- Examples:
  - Al_2O_3, aluminum oxide
  - $Ba(OH)_2$, barium hydroxide
  - KCl, potassium chloride
  - ZnS, zinc sulfide

Naming Binary Ionic Compounds

- Numerical prefixes are never used.
- If the metal can exist in more oxidation states, its oxidation number is included to the name
- Examples:
  - $FeCl_3$, iron(III) chloride (ferric chloride)
  - $FeCl_2$, iron(II) chloride (ferrous chloride)
  - CuO, copper(II) oxide (cupric oxide)
  - Cu_2O, copper(I) oxide (cuprous oxide)

Naming Binary Molecular Compounds

- Compounds of two nonmetals (e.g. oxides of nonmetals)
- Name of less electronegative element + stem of the other element + **-ide**
- Numerical prefixes precede names of both nonmetals
- Examples:
  - CO, carbon monoxide
  - N_2O_5, dinitrogen pentoxide
  - CCl_4, carbon tetrachloride
  - H_2S, hydrogen sulfide

Numerical prefixes:

Number	Prefix
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
8	nona-
10	deca-

Naming acids and their salts

Brønsted-Lowry concept of acids and bases:

- Acid is a proton donor
- Base is a proton acceptor

$$HCl(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + Cl^-(aq)$$

Acid — Base — Conjugate acid — Conjugate base

Salt: ionic compound, product of neutralization reaction between acid and base (the acidic proton replaced with metal cation)

Naming acids and their salts

A) hydroacids:

Gaseous nonmetallic hydrides whose aqueous solutions are acidic

E.g. HCl, hydrochloric acid (aqueous hydrogen chloride), salts: chloride

Likewise:

- HBr, hydrobromic acid, salts bromides
- H_2S, hydrosulfuric acid, salts sulfides
- HCN, hydrocyanic acid, salts cyanides

Naming acids and their salts

B) oxo-acids:

Central atom + -OH groups, protons dissociate from oxygen. In salts appear as polyatomic anions

If only one oxidation state of the central atom is possible:

Stem of the central atom + **-ic acid**

E.g. H_2CO_3, carbonic acid, salt: carbonate

Naming acids and their salts

B) oxo-acids:

If there are two possible oxidation states of the central atom:

Higher ox. number: **-ic acid**, salt: **-ate**

Lower ox. number: **-ous acid**, salt: **-ite**

Example:

H_2SO_4, sulfuric acid, salt: sulfate

H_2SO_3, sulfurous acid, salt: sulfite

Naming acids and their salts

B) oxo-acids:

If there are more than two possible oxidation states of the central atom, prefixes are used:

HClO, hypochlorous acid, salt: hypochlorite

$HClO_2$, chlorous acid, salt: chlorite

$HClO_3$, chloric acid, salt: chlorate

$HClO_4$, perchloric acid, salt: perchlorate

The oxidation numbers can also be used with metals in anions:

MnO_4^{2-}: **manganate(VI)** or just **manganate**

MnO_4^{-}: **manganate(VII)** or **permanganate**

$[Fe(CN)_6]^{4-}$: **hexacyanoferrate(II)** or **ferrocyanide**

$[Fe(CN)_6]^{3-}$: **hexacyanoferrate(III)** or **ferricyanide**

Naming coordination compounds

- Names of neutral ligands:
  - H_2O aqua
  - NH_3 ammin
  - NO nitrosyl
  - CO carbonyl
- Names of anionic ligands always end in –o:
  - F^- fluorido
  - Cl^- chlorido
  - Br^- bromido
  - I^- iodido
  - OH^- hydroxido
  - CN^- cyanido
  - etc..

Naming coordination compounds

1. **Complex particle is cation:**

   e.g. $[Cu(NH_3)_4]SO_4$

   $[Cu(NH_3)_4]^{2+}$ + SO_4^{2-}

   Tetraamminecopper(II) sulfate

2. **Complex particle is anion:**

   e.g. $K_3[CoF_6]$

   3 K^+ + $[CoF_6]^{3-}$

   Potassium hexafluoridocobaltate(III)

Naming coordination compounds

3. **Both cation and anion are complexes:**

   e.g. $[Pt(NH_3)_4][PtCl_4]$

   $[Pt(NH_3)_4]^{2+}$ + $[PtCl_4]^{2-}$

   Tetraammineplatinum(II) tetrachloridoplatinate(II)

4. **Neutral complexes:**

   e.g. $[CrCl_3(H_2O)_3]$

   Triaquatrichloridochromium(III) complex

Additive IUPAC names of oxo-acids

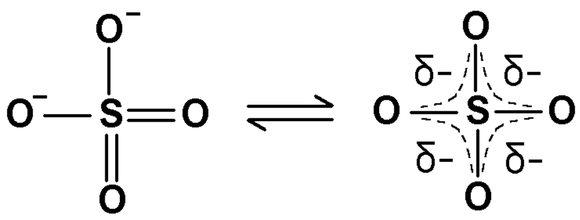
E.g.

Sulfuric acid H_2SO_4:

– Dihydroxidodioxidosulfur

Sulfate SO_4^{2-}:

– Tetraoxidosulfate(2-)



Ionic equations

Reaction I

Stoichiometric equation:

$AgNO_3 + HCl \longrightarrow AgCl + HNO_3$

Reaction I

Stoichiometric equation:

$AgNO_3 + HCl \longrightarrow AgCl + HNO_3$

Ionic equation:

$Ag^+ + NO_3^- + H^+ + Cl^- \longrightarrow AgCl + H^+ + NO_3^-$

Net ionic equation:

$Ag^+ + \cancel{NO_3^-} + \cancel{H^+} + Cl^- \longrightarrow AgCl + \cancel{H^+} + \cancel{NO_3^-}$

$\boldsymbol{Ag^+ + Cl^- \longrightarrow AgCl}$ white ppt

Also possible:

$AgNO_3(aq) + HCl(aq) \longrightarrow AgCl(s) + HNO_3(aq)$

$$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$$

(aq)	... aqueous
(s)	... solid
(l)	... liquid
(g)	... gaseous

What combinations of cations and anions are insoluble?

- All nitrates (NO_3^-) and acetates (CH_3COO^-) are soluble
- All salts of Na, K, Li, and NH_4^+ are soluble
- All chlorides, bromides and iodides are soluble except salts of Pb^{2+}, Ag^+, and Hg_2^{2+}
- Most sulfate salts are soluble except $BaSO_4$, $PbSO_4$, $HgSO_4$, and $CaSO_4$.
- Most metal hydroxides are insoluble. Soluble are only LiOH, NaOH, KOH, $Ba(OH)_2$ and $Ca(OH)_2$.
- Most sulfides (S^{2-}), carbonates (CO_3^{2-}) and phosphates (PO_4^{3-}) are insoluble.

Reaction II

$2\,NaOH + (NH_4)_2SO_4 \longrightarrow 2\,NH_3 + Na_2SO_4 + 2H_2O$

ionic:

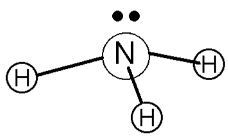
$2Na^+ + 2OH^- + 2NH_4^+ + SO_4^{2-} \longrightarrow 2NH_3 + 2Na^+ + SO_4^{2-} + 2H_2O$

net ionic:

$\cancel{2}\cancel{Na^+} + \cancel{2}OH^- + \cancel{2}NH_4^+ + \cancel{SO_4^{2-}} \longrightarrow \cancel{2}NH_3 + \cancel{2Na^+} + \cancel{SO_4^{2-}} + \cancel{2}H_2O$

$$\mathbf{OH^- + NH_4^+ \longrightarrow NH_3 + H_2O}$$

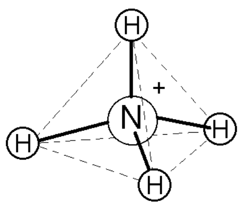
Ammonia gas: NH_3, $NH_3(g)$



Aqueous ammonia: $NH_3(aq)$, $NH_3.H_2O$, NH_4OH

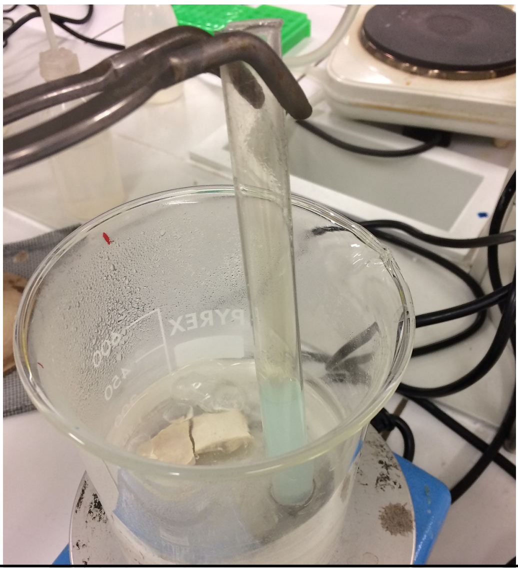
$\mathbf{NH_3.H_2O \longleftrightarrow NH_4OH}$

$\mathbf{NH_4OH \longleftrightarrow NH_4^+ + OH^-}$



Reaction III

$4\ HNO_3(aq) + Cu(s) \longrightarrow Cu(NO_3)_2 + 2\ NO_2(g) + 2\ H_2O(l)$



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$4\ HNO_3(aq) + Cu(s) \longrightarrow Cu(NO_3)_2 + 2\ NO_2(g) + 2\ H_2O(l)$

Ionic:

$4\ H^+ + 4\ NO_3^- + Cu \longrightarrow Cu^{2+} + 2\ NO_3^- + 2\ NO_2 + 2\ H_2O$

Net ionic:

$4\ H^+ + \not{4}\ NO_3^- + Cu \longrightarrow Cu^{2+} + 2\ \not{NO_3^-} + 2\ NO_2 + 2\ H_2O$

$4\ H^+ + 2\ NO_3^- + Cu \longrightarrow Cu^{2+} + 2\ NO_2 + 2\ H_2O$

Reaction III

$4\ HNO_3(aq) + Cu(s) \longrightarrow Cu(NO_3)_2 + 2\ NO_2(g) + 2\ H_2O(l)$

Ionic:

$4\ H^+ + 4\ NO_3^- + Cu \longrightarrow Cu^{2+} + 2\ NO_3^- + 2\ NO_2 + 2\ H_2O$

Oxidation numbers:

H: +I → +I
Cu: 0 → +II (loss of 2 e^- …oxidized)
N: +V → +V, + IV (gain of + e^- …reduced)
O: –II → –II

Writing Ionic equations: Summary

1. write correct and balanced stoichiometric equation first
2. rewrite to ionic: write separately any species that exist separately and indicate its charge if present, but write together what exists joined (usually a precipitate of insoluble salt, or a soluble coordination complex)
3. Cancel out all species not involved in the reaction
4. Check that the equation is still balanced