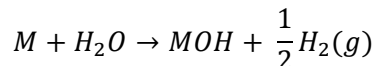


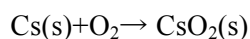
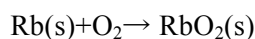
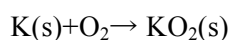
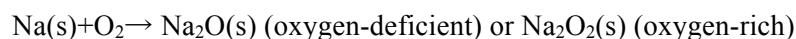
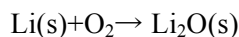
**Reactions of s-block elements with water, oxygen, and halogens, their color in flame tests**

Group 1 (Alkali Metals): let M be a metal of Group 1 of the Periodic Table



MOH dissociates completely to form a strongly basic solution. Li is the strongest s-block reducing agent, due to a large energy of hydration of small  $Li^+$  ions. However, it reacts more slowly than Na or K, which melt upon reaction with water due to the immense amount of heat energy released, thereby increasing the surface area available for reaction.

Reactions with  $O_2(g)$ :



$Na_2O_2(s) + 2H_2O \rightarrow 2Na^+(aq) + H_2O_2(aq) + 2OH^-$  (hydrogen peroxide is a very strong oxidizing agent – stronger than even chlorine!)

Potassium, Rubidium, and Cesium superoxides react with water to produce hydroxide ion, oxygen gas, and hydrogen peroxide. They react with carbon dioxide to form a metal carbonate and oxygen gas. The production of oxygen gas makes superoxides very useful in self-contained breathing apparatuses.

If X is a halogen, the alkali metals react with  $X_2$  to form MX (s).

Group 2 (Alkaline Earth Metals): let M be a metal of Group 2 of the Period Table

MO oxides are basic:  $MO(s) + H_2O(l) \rightarrow M^{2+} + 2OH^-$

M as an element is also basic:  $M(s) + 2H_2O \rightarrow M^{2+} + 2OH^- + H_2(g)$

Ca, Sr, and Ba react vigorously with water at 298K. Be and Mg are less easily oxidized, and show no observable reaction with water at 298K, although Mg reacts with boiling water.

BeO and  $Be(OH)_2$  are amphoteric, mostly due to a very strong covalent character. All alkaline earth metals are prepared/purified by electrolysis of their molten metal chloride ( $MCl_2$ ).

$Mg(s) + O_2(g) \rightarrow MgO(s) + \text{light energy}$  (Mg(s) is also a useful structural material.)

If X is a halogen, the alkaline earth metals react with  $X_2(g)$  to form  $MX_2(s)$ .

Group 2 hydrides will only form when M=Ca, Sr, or Ba at high temp, or Mg at high pressure. They are very strongly covalently bonded in nature (not ionically bonded).

Regarding the formation of oxides, all of the alkaline earth metals react with oxygen to form an oxide in a 1:1 ratio (MO), although Ba can form barium superoxide in the presence of excess oxygen.

Element	Flame Color
Li	Red
Na	Strong and Persistent Yellow/Orange
K	Lilac (Pink)
Rb	Red (Red-Violet)
Cs	Blue/Violet
Ca	Orange-Red
Sr	Red
Ba	Pale Green
Mg	Very, very White

### Stoichiometry, reactions, and properties of binary non-metal hydrides

Elements in group 13 to 17 (p-block) form covalent hydrides (or **nonmetal hydrides**). In group 12 zinc hydride is a common chemical reagent but cadmium hydride and mercury hydride are very unstable and esoteric. In group 13 boron hydrides exist as a highly reactive monomer  $\text{BH}_3$ , as an adduct for example ammonia borane or as dimeric diborane and as a whole group of BH cluster compounds. Alane ( $\text{AlH}_3$ ) is a polymer. Gallium exists as the dimer digallane. Indium hydride is only stable below  $-90\text{ }^\circ\text{C}$  ( $-130\text{ }^\circ\text{F}$ ).

In group 14 the total number of possible binary saturated compounds with carbon of the type  $\text{C}_n\text{H}_{2n+2}$  is very large. Going down the group the number of binary silicon compounds (silanes) is small (straight or branched but rarely cyclic) for example disilane and trisilane. For germanium only 5 linear chain binary compounds are known as gases or volatile liquids. Examples are n-pentagermane, isopentagermane and neopentagermane. Of tin only the distannane is known. Plumbane is an unstable gas.

Non-classical hydrides are those in which extra hydrogen molecules are coordinated as a ligand on the central atoms. These are very unstable but some have been shown to exist.

### The periodic table of the stable binary hydrides

The relative stability of binary hydrogen compounds and alloys at standard temperature and pressure can be inferred from their standard enthalpy of formation values.

$\text{H}_2$ 0						He
LiH -91	BeH <sub>2</sub> 125					Ne
Na	MgH <sub>2</sub>	AlH <sub>3</sub>	SiH <sub>4</sub> - 74.8	NH <sub>3</sub> - 46.8	H <sub>2</sub> O - 243	Ar
				PH <sub>3</sub> 5.	H <sub>2</sub> S -	
					HCl	

H - 57	-75												31	4	207	-93	
KH -58	CaH <sub>2</sub> - 174	Sc H <sub>3</sub>	TiH 1.7	VH	Cr H	Mn	Fe	Co	Ni	Cu H	Zn H <sub>2</sub>	GaH 3	GeH <sub>4</sub> 92	AsH <sub>3</sub> 67	H <sub>2</sub> Se 30	HBr -36.5	Kr
Rb H - 47	SrH <sub>2</sub> - 177	YH 3	ZrH 2	Nb H <sub>x</sub>	Mo	Tc	Ru	Rh	Pd H	Ag	Cd H <sub>2</sub>	InH <sub>3</sub>	SnH <sub>4</sub> 163	SbH <sub>3</sub> 146	H <sub>2</sub> Te 100	HI 2 6.6	Xe
CsH -50	BaH <sub>2</sub> - 172		Hf H <sub>2</sub>	Ta H	W	Re	Os	Ir	Pt	Au	Hg H	Tl	PbH <sub>4</sub> 252	BiH <sub>3</sub> 2 77	H <sub>2</sub> Po	At	Rn
Fr	Ra		Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Uut	Uuq	Uup	Uuh	Uus	Uuo

↓

La H <sub>2</sub>	Ce H <sub>2</sub>	PrH 2	Nd H <sub>2</sub>	Pm H <sub>2</sub>	Sm H <sub>2</sub>	Eu H <sub>2</sub>	Gd H <sub>2</sub>	Tb H <sub>2</sub>	Dy H <sub>2</sub>	HoH 2	ErH <sub>2</sub>	TmH <sub>2</sub>	YbH <sub>2</sub>	LuH 2	
Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr	

### Binary compounds of hydrogen

Covalent hydrides	metallic hydrides.
Ionic hydrides	Intermediate hydrides.
Do not exist	Not assessed

## Molecular hydrides

Most monomeric hydrides are isolable only under extreme conditions (i.e. at cryogenic temperatures, and often embedded in a rare gas matrix). This is generally attributable to poor contribution of the atomic orbitals of the respective atoms with the s-orbital of hydrogen; and to the low activation enthalpies of autopolymerisation reactions, which electron-deficient monomers are prone to undergo. The table below shows the monomeric hydride for each element, which is closest to, but not surpassing its heuristic valence. A heuristic valence is the valence of an element that strictly obeys the octet, duodecet, and other valence rules. Where available, both the enthalpy of formation for each monomer and the enthalpy of formation for the hydride in its standard state is shown (in brackets) to give a rough indication of which monomers tend to undergo aggregation to lower enthalpic states. For example, monomeric lithium hydride has an enthalpy of formation of 139 kJ mol<sup>-1</sup>, whereas solid lithium hydride has an enthalpy of -91 kJ mol<sup>-1</sup>. This means that it is energetically favourable for a mole of monomeric LiH to aggregate into the ionic solid, losing 230 kJ as a consequence. Aggregation can occur as a chemical association, such as polymerisation, or it can occur as an electrostatic association, such as the formation of hydrogen-bonding in water.

## Classical hydrides

### Classical hydrides

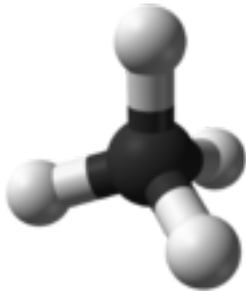
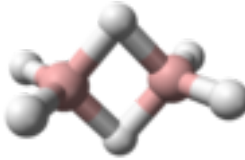
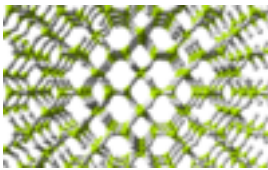
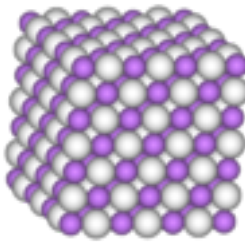
1	2	3	4	5	6	5	4	3	2	1	2	3	4	3	2	1
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<b>H</b> 2 0																				
<b>LiH</b> <sup>[7]</sup> 1 39 (-91)	<b>Be</b> <b>H</b> 2 <sup>[8]</sup> <sub>12</sub> 3											<b>BH</b> 3 <sup>[9]</sup> 1 07 (41)	<b>C</b> <b>H</b> 4-7 5	<b>NH</b> 3-46	<b>H</b> 2O <sup>-2</sup> 42 (-286 )	<b>HF</b> <sup>-2</sup> 73				
<b>NaH</b> <sup>[10]</sup> 140 (-56)	<b>Mg</b> <b>H</b> 2 14 2 (-76 )											<b>AlH</b> 3 <sup>[11]</sup> 1 23 (-46)	<b>Si</b> <b>H</b> 4 <sub>34</sub>	<b>PH</b> 3 5	<b>H</b> 2S <sup>-21</sup>	<b>HCl</b> <sup>-</sup> 92				
<b>KH</b> 132 (-58)	<b>Ca</b> <b>H</b> 2 19 2 (-17 4)	<b>Sc</b> <b>H</b> 3	<b>Ti</b> <b>H</b> 4	<b>V</b> <b>H</b> 2 <sup>[12]</sup>	<b>Cr</b> <b>H</b> 2 <sup>[13]</sup>	<b>Mn</b> <b>H</b> 2 <sup>[14]</sup>	<b>Fe</b> <b>H</b> 2 <sup>[15]</sup> <sub>3</sub> 24	<b>Co</b> <b>H</b> 2 <sup>[16]</sup>	<b>NiH</b> 2 <sup>[17]</sup>	<b>CuH</b> <sup>[18]</sup> 278 (28)	<b>Zn</b> <b>H</b> 2 <sup>[19]</sup> <sub>1</sub> 62	<b>GaH</b> 3 <sup>[20]</sup> <sub>15</sub> 1	<b>Ge</b> <b>H</b> 4 <sub>92</sub>	<b>As</b> <b>H</b> 3 <sub>67</sub>	<b>H</b> 2Se <sub>30</sub>	<b>HBr</b> -36				
<b>RbH</b> 13 2 (-47)	<b>SrH</b> 2 20 1 (-17 7)	<b>Y</b> <b>H</b> 3	<b>Zr</b> <b>H</b> 4	<b>Nb</b> <b>H</b> 4 <sup>[12]</sup>	<b>Mo</b> <b>H</b> 6 <sup>[21]</sup>	Tc	<b>Ru</b> <b>H</b> 2 <sup>[15]</sup>	<b>Rh</b> <b>H</b> 2 <sup>[22]</sup>	<b>PdH</b> <sup>[23]</sup> 361	<b>AgH</b> <sup>[18]</sup> 288	<b>Cd</b> <b>H</b> 2 <sup>[19]</sup> <sub>1</sub> 83	<b>InH</b> 3 <sup>[24]</sup> <sub>22</sub> 2	<b>Sn</b> <b>H</b> 4 <sub>16</sub> 3	<b>Sb</b> <b>H</b> 3 <sub>146</sub>	<b>H</b> 2Te <sub>10</sub> 0	<b>HI</b> 27				
<b>CsH</b> 11 9 (-50)	<b>Ba</b> <b>H</b> 2 21 3 (-17 7)		<b>Hf</b> <b>H</b> 4	<b>Ta</b> <b>H</b> 4 <sup>[12]</sup>	<b>WH</b> 6 <sup>[25]</sup> <sub>5</sub> 86	<b>Re</b> <b>H</b> 4 <sup>[14]</sup>	<b>Os</b>	<b>Ir</b>	<b>PtH</b> 2 <sup>[26]</sup>	<b>AuH</b> <sup>[18]</sup> 295	<b>Hg</b> <b>H</b> 2 <sup>[27]</sup> <sub>1</sub> 01	<b>TlH</b> 3 <sup>[28]</sup> <sub>29</sub> 3	<b>Pb</b> <b>H</b> 4 <sub>25</sub> 2	<b>Bi</b> <b>H</b> 3 <sub>247</sub>	<b>H</b> 2Po <sub>16</sub> 7	<b>HAt</b> <sub>8</sub> 8				
Fr	Ra		Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Uut	Fl	Uup	Lv	Uus				

↓

3	4	5	6	7	8	7	6	5	4	3	2	1	2	3
<b>La</b> <b>H</b> 3	<b>Ce</b> <b>H</b> 4	<b>Pr</b> <b>H</b> 3	<b>Nd</b> <b>H</b> 4	Pm	<b>Sm</b> <b>H</b> 4	<b>Eu</b> <b>H</b> 2 <sup>[29]</sup>	<b>GdH</b> 3	<b>TbH</b> 3	<b>Dy</b> <b>H</b> 4	<b>HoH</b> 3	<b>Er</b> <b>H</b> 2	<b>Tm</b> <b>H</b>	<b>YbH</b> 2	<b>LuH</b> 3
Ac	<b>Th</b>	<b>Pa</b>	<b>UH</b>	<b>Np</b>	<b>Pu</b>	A	<b>Cm</b>	<b>Bk</b>	<b>Cf</b>	<b>Es</b>	<b>Fm</b>	<b>Md</b>	<b>No</b>	<b>Lr</b>

H	4 <sup>[31]</sup>	m																	
4 <sup>[30]</sup>																			

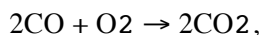
Legend			
Monomeric covalent hydride		Oligomeric covalent hydride	
Polymeric covalent hydride		Ionic hydride	

This table includes the thermally unstable dihydrogen complexes for the sake of completeness. As with the above table, only the complexes with the most complete valence is shown, to the negligence of the most stable complex.

### Common reactions of carbon, nitrogen and sulfur oxides (CO, CO<sub>2</sub>, NO, NO<sub>2</sub>, N<sub>2</sub>O<sub>4</sub>, SO<sub>2</sub>, SO<sub>3</sub>)

#### • *Carbon monoxide.*

Carbon monoxide is produced when graphite (one of the naturally occurring forms of elemental carbon) is heated or burned in a limited amount of oxygen. The reaction of steam with red-hot coke also produces carbon monoxide along with hydrogen gas (H<sub>2</sub>). (Coke is the impure carbon residue resulting from the burning of coal.) This mixture of CO and H<sub>2</sub> is called water gas and is used as an industrial fuel. In the laboratory, carbon monoxide is prepared by heating formic acid, HCOOH, or oxalic acid, H<sub>2</sub>C<sub>2</sub>O<sub>4</sub>, with concentrated sulfuric acid, H<sub>2</sub>SO<sub>4</sub>. The sulfuric acid removes the elements of water (i.e., H<sub>2</sub>O) from the formic or oxalic acid and absorbs the water produced. Because carbon monoxide burns readily in oxygen to produce carbon dioxide,

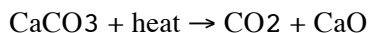


it is useful as a gaseous fuel. It is also useful as a metallurgical reducing agent because at high temperatures it reduces many metal oxides to the elemental metal. For example, copper(II) oxide, CuO, and iron(III) oxide, Fe<sub>2</sub>O<sub>3</sub>, are both reduced to the metal by carbon monoxide.

Carbon monoxide is an extremely dangerous poison. Because it is an odourless and tasteless gas, it gives no warning of its presence. It binds to the hemoglobin in blood to form a compound that is so stable that it cannot be broken down by body processes. When the hemoglobin is combined with carbon monoxide, it cannot combine with oxygen; this destroys the ability of hemoglobin to carry essential oxygen to all parts of the body. Suffocation can occur if sufficient amounts of carbon monoxide are present to form complexes with the hemoglobin.

#### • *Carbon dioxide.*

Carbon dioxide is produced when any form of carbon or almost any carbon compound is burned in an excess of oxygen. Many metal carbonates liberate CO<sub>2</sub> when they are heated. For example, calcium carbonate (CaCO<sub>3</sub>) produces carbon dioxide and calcium oxide (CaO).

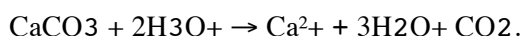


The fermentation of glucose (a sugar) during the preparation of ethanol, the alcohol found in beverages such as beer and wine, produces large quantities of CO<sub>2</sub> as a by-product.



glucoseethanol

In the laboratory CO<sub>2</sub> can be prepared by adding a metal carbonate to an aqueous acid; e.g.,



Carbon dioxide is a colourless and essentially odourless gas that is 1.5 times as dense as air. It is not toxic, although a large concentration could result in suffocation simply by causing a lack of oxygen in the body. All carbonated beverages contain dissolved CO<sub>2</sub>; hence the name carbonated. One litre (1.06 quarts) of water at 20° C dissolves 0.9 litre of CO<sub>2</sub> at one atmosphere, forming carbonic acid (H<sub>2</sub>CO<sub>3</sub>, discussed below in the section Oxyacids and their salts), which has a mildly acidic (sour) taste. Solid CO<sub>2</sub> sublimates at normal atmospheric pressure. Thus, solid CO<sub>2</sub>, called dry ice, is a valuable refrigerant that is always free of the liquid form. Carbon dioxide is also used as a fire extinguisher, because most substances do not burn in it, and it is readily available and inexpensive. Air containing as little as 2.5 percent CO<sub>2</sub> extinguishes a flame.

In the absence of oxygen, CO<sub>2</sub> is very strong oxidizing agent. It reacts with metals to form solid carbon and a metal carbonate.

In excess oxygen, carbon burns to form carbon dioxide. In oxygen-deficient settings, carbon instead burns to form carbon monoxide, which can later react with oxygen gas to form carbon dioxide.

### ***Binary nitrogen oxide compounds.***

$\text{NO} + \text{O}_2 \rightarrow \text{NO}_2$  (NO can also react with partly oxidized organic species in atmosphere to produce nitrogen dioxide.)

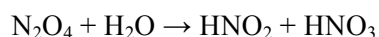
$3\text{NO}_2 + \text{H}_2\text{O} \rightarrow 2\text{HNO}_3 + \text{NO}$  (formation of acid rain)

$\text{NO} + \text{X}_2 \rightarrow \text{NOX}$  (X is Cl, F, or Br)

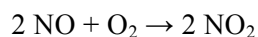
$\text{NH}_3 + \text{O}_2 \rightarrow \text{NO}$  (Pt heterogeneous catalyst)

HNO<sub>3</sub>, concentrated yields NO<sub>2</sub>+H<sub>2</sub>O when reacted with metals. HNO<sub>3</sub>, dilute (e.g. 6M) yields NO+H<sub>2</sub>O.

Dinitrogen tetraoxide (N<sub>2</sub>O<sub>4</sub>) forms an equilibrium mixture with NO<sub>2</sub>. It is also an intermediate in the manufacture of nitric acid, and is used in the synthesis of metal nitrates.



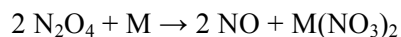
The coproduct HNO<sub>2</sub> upon heating disproportionates to NO and more nitric acid. When exposed to oxygen, NO is converted back into nitrogen dioxide:



The resulting  $\text{NO}_2$  (and  $\text{N}_2\text{O}_4$ , obviously) can be returned to the cycle to give the mixture of nitrous and nitric acids again.

### Synthesis of metal nitrates

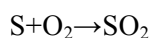
$\text{N}_2\text{O}_4$  behaves as the salt  $[\text{NO}^+][\text{NO}_3^-]$ , the former being a strong oxidant:



where  $\text{M} = \text{Cu}, \text{Zn}, \text{or Sn}$ .

$\text{NO}$  and  $\text{NO}_2$  are stable when isolated but fairly chemically reactive ( $\text{NO}_2$  can combust with hydrocarbons).  $\text{NO}$  is colorless but  $\text{NO}_2$  is a brownish red.  $\text{N}_2\text{O}$  are stable when isolated and not really chemically reactive (could be because asymmetry produces to high of an activation energy for a reaction involving nitrous oxide to be feasible).

### Binary sulfur-oxide compounds



$\text{SO}_2$  and  $\text{SO}_3$  are the acid anhydrides of  $\text{H}_2\text{SO}_3$  and  $\text{H}_2\text{SO}_4$ , respectively. This thus makes the gases good reducing agents.

Combustion of metallic sulfides (e.g. pyrite  $\text{FeS}_2$ ) yields  $\text{SO}_2$ . Sulfur dioxide has a pungent odor.

### Common oxidation states of p-block elements, stoichiometry of common halides and oxoacids

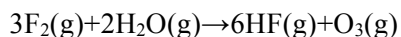
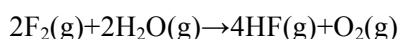


Nitrogen, phosphorus, and chlorine have very, very wide ranges of oxidation states. Fluorine will only exhibit -1 oxidation number. Oxidation states of larger p-block elements follow are roughly the same as their group "leader" in  $n=2$  period, although some may exhibit a wider range as one goes down the group (e.g. halides).

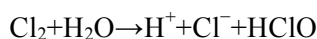
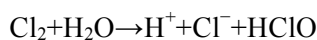
### Reaction of halogens with water

From a standard reduction potential table, it is determined that iodine and bromine cannot oxidize water to oxygen because they have smaller reduction potentials than oxygen. Thus, iodine and bromine do not react with water. However, fluorine and chlorine have larger reduction potentials, and can oxidize water.

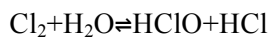
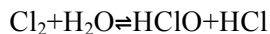
Fluorine reacts with water vapor to form oxygen and ozone:



The reaction of water with chlorine, shown below, proceeds very slowly.



Chlorine and bromine are moderately soluble in water. These solutions form solid hydrates within an ice lattice. These solutions are good oxidizing agents. Chlorine reacts reversibly with water to produce acids as in the following example, in which chloric acid and hydrochloric acid are formed:



Iodine is slightly soluble in water. It has the lowest standard reduction potential of the halogens, and is therefore the least powerful oxidizing agent. Air and other reagents can oxidize acidified solution of iodide ions.

### Common oxidation states of first row transition metals

Element Name and Symbol	Atomic Number	Common Oxidation States	Electron Configuration	
Scandium (Sc)	21	+3	Sc: [Ar] 4s <sup>2</sup> 3d <sup>1</sup>	Sc: [Ar] $\frac{\uparrow\downarrow}{4s}$ $\frac{1}{\text{---}}$ $\frac{\text{---}}{\text{---}}$ $\frac{\text{---}}{\text{---}}$ $\frac{\text{---}}{\text{---}}$ 3d
Titanium (Ti)	22	+4	Ti: [Ar] 4s <sup>2</sup> 3d <sup>2</sup>	Ti: [Ar] $\frac{\uparrow\downarrow}{4s}$ $\frac{1}{\text{---}}$ $\frac{1}{\text{---}}$ $\frac{\text{---}}{\text{---}}$ $\frac{\text{---}}{\text{---}}$ 3d
Vanadium (V)	23	+2, +3, +4, +5	V: [Ar] 4s <sup>2</sup> 3d <sup>3</sup>	V: [Ar] $\frac{\uparrow\downarrow}{4s}$ $\frac{1}{\text{---}}$ $\frac{1}{\text{---}}$ $\frac{1}{\text{---}}$ $\frac{\text{---}}{\text{---}}$ $\frac{\text{---}}{\text{---}}$ 3d
Chromium (Cr)	24	+2, +3, +6	Cr: [Ar] 4s <sup>1</sup> 3d <sup>5</sup>	Cr: [Ar] $\frac{1}{4s}$ $\frac{1}{\text{---}}$ $\frac{1}{\text{---}}$ $\frac{1}{\text{---}}$ $\frac{1}{\text{---}}$ $\frac{1}{\text{---}}$ 3d
Manganese (Mn)	25	+2, +3, +4, +6, +7	Mn: [Ar] 4s <sup>2</sup> 3d <sup>5</sup>	Mn: [Ar] $\frac{\uparrow\downarrow}{4s}$ $\frac{1}{\text{---}}$ $\frac{1}{\text{---}}$ $\frac{1}{\text{---}}$ $\frac{1}{\text{---}}$ $\frac{1}{\text{---}}$ 3d
Iron (Fe)	26	+2, +3	Fe: [Ar] 4s <sup>2</sup> 3d <sup>6</sup>	Fe: [Ar] $\frac{\uparrow\downarrow}{4s}$ $\frac{\uparrow\downarrow}{\text{---}}$ $\frac{1}{\text{---}}$ $\frac{1}{\text{---}}$ $\frac{1}{\text{---}}$ $\frac{1}{\text{---}}$ 3d
Cobalt (Co)	27	+2, +3	Co: [Ar] 4s <sup>2</sup> 3d <sup>7</sup>	Co: [Ar] $\frac{\uparrow\downarrow}{4s}$ $\frac{\uparrow\downarrow}{\text{---}}$ $\frac{\uparrow\downarrow}{\text{---}}$ $\frac{1}{\text{---}}$ $\frac{1}{\text{---}}$ $\frac{1}{\text{---}}$ 3d
Nickel (Ni)	28	+2	Ni: [Ar] 4s <sup>2</sup> 3d <sup>8</sup>	Ni: [Ar] $\frac{\uparrow\downarrow}{4s}$ $\frac{\uparrow\downarrow}{\text{---}}$ $\frac{\uparrow\downarrow}{\text{---}}$ $\frac{\uparrow\downarrow}{\text{---}}$ $\frac{1}{\text{---}}$ $\frac{1}{\text{---}}$ 3d
Copper (Cu)	29	+1, +2	Cu: [Ar] 4s <sup>1</sup> 3d <sup>10</sup>	Cu: [Ar] $\frac{1}{4s}$ $\frac{\uparrow\downarrow}{\text{---}}$ $\frac{\uparrow\downarrow}{\text{---}}$ $\frac{\uparrow\downarrow}{\text{---}}$ $\frac{\uparrow\downarrow}{\text{---}}$ $\frac{\uparrow\downarrow}{\text{---}}$ 3d
Zinc (Zn)	30	+2	Zn: [Ar] 4s <sup>2</sup> 3d <sup>10</sup>	Zn: [Ar] $\frac{\uparrow\downarrow}{4s}$ $\frac{\uparrow\downarrow}{\text{---}}$ $\frac{\uparrow\downarrow}{\text{---}}$ $\frac{\uparrow\downarrow}{\text{---}}$ $\frac{\uparrow\downarrow}{\text{---}}$ $\frac{\uparrow\downarrow}{\text{---}}$ 3d

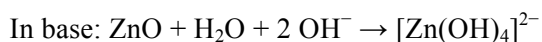
Ion	Color in Solution
Dichromate	Orange
Chromate	Yellow
Chromium(III)	Green
Chromium(II)	Blue
Scandium(III), Zinc(II), Copper(I), Titanium(IV), Silver (I)	No color
Vanadium(II)	Violet
Vanadium(III)	Green
Vanadium(IV)	Blue
Vanadium(V)	Yellow
Manganese(II)	Very pale pink
Manganese(IV)	Brown
Manganese(VII) (Permanganate)	Purple
Iron(II)	Very pale green



Iron(III)	Dark yellowish brown
Cobalt(II)	Red
Nickel(II)	Green
Copper(II)	Blue/Green

### Dissolution of these metals and Al, amphoteric hydroxides

Zinc oxide (ZnO) reacts with both acids and with bases:



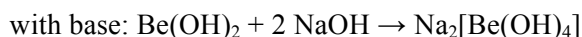
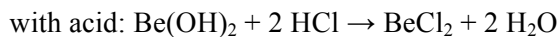
This reactivity can be used to separate different cations, such as zinc(II), which dissolves in base, from manganese(II), which does not dissolve in base.

Aluminium hydroxide is also amphoteric: (Chromium hydroxide acts in a very similar fashion – just substitute Cr for Al in reactions given)

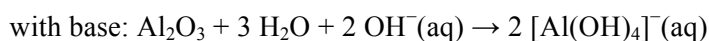
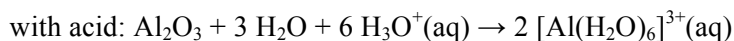


Some other amphoteric compounds include:

Beryllium hydroxide (also beryllium oxide)



Aluminium oxide



Lead(II) oxide



Some other elements which form amphoteric oxides are gallium ( $\text{Ga}_2\text{O}_3$ ), indium, scandium, titanium, zirconium, vanadium, chromium, iron, cobalt, copper, silver, gold, germanium, tin, antimony, bismuth, and tellurium.

### Permanganate, chromate, dichromate ions and their redox reactions

Ion	In Acidic Solution	In Neutral/Basic Solution
$\text{MnO}_4^-$	$\text{Mn}^{2+}$	$\text{MnO}_2$
$\text{CrO}_4^{2-}$	$\text{Cr}^{3+}$	$\text{CrO}_2^-$
$\text{Cr}_2\text{O}_7^{2-}$	$2\text{Cr}^{3+}$	$\text{Cr}(\text{OH})_3$

Chromate and Dichromate are better oxidizing agents in acidic, rather than alkaline, conditions.

$\text{Cr}_2\text{O}_7^{2-}$  reacts with a base to form  $\text{CrO}_4^{2-} + \text{H}_2\text{O}$

### Iodometry (reaction of thiosulphate and iodine)

titration based on redox reaction  $\text{S}_2\text{O}_3^{2-} + \text{I}_2 \rightarrow \text{S}_4\text{O}_6^{2-} + 2\text{I}^-$  (unbalanced)

$\text{I}_2$  forms a complex with starch that yields a dark blue color.  $2\text{I}^-$  with starch is colorless. (starch serves as an indicator)

Pretty much all other reactants will react with thiosulfate to produce sulfate instead of  $\text{S}_4\text{O}_6^{2-}$

### Identification of $\text{Ag}^+$ , $\text{Ba}^{2+}$ , $\text{Fe}^{3+}$ , $\text{Cu}^{2+}$ , $\text{Cl}^-$ , $\text{CO}_3^{2-}$ , $\text{SO}_4^{2-}$

Qualitative Analysis of Inorganic Ions			
Group Number	Solution Tested	Precipitating Agent	Precipitated Compound(s)
1	Unknown	0.1 M HCl	$\text{PbCl}_2$ , $\text{AgCl}$ , $\text{Hg}_2\text{Cl}_2$
2	Filtrate/Supernatant from Group 1	$\text{H}_2\text{S}$ at pH 1	$\text{HgS}$ , $\text{PbS}$ , $\text{CuS}$ , $\text{CdS}$ , $\text{Bi}_2\text{S}_3$ , $\text{As}_2\text{S}_3$ , $\text{Sb}_2\text{S}_3$ , $\text{SnS}_2$
3	Filtrate/Supernatant from Group 2	$\text{H}_2\text{S}$ at pH 10	$\text{MnS}$ , $\text{FeS}$ , $\text{NiS}$ , $\text{CoS}$ , $\text{ZnS}$ , $\text{Fe}(\text{OH})_3$ , $\text{Al}(\text{OH})_3$ , $\text{Cr}(\text{OH})_3$ ,
4	Filtrate/Supernatant from Group 3	$\text{CO}_3^{2-}$ at pH 10	$\text{MgCO}_3$ , $\text{CaCO}_3$ , $\text{SrCO}_3$ , $\text{BaCO}_3$ ,
5	Unknown	None	Soluble ions $\text{Na}^+$ , $\text{K}^+$ , $\text{NH}_4^+$

Start with unknown solution. Add precipitating agent. Remove precipitates by filtration/centrifugation. Move to next group. After each group has precipitate, additional separations and conformation tests are run to identify individual ions. e.g. Group 1 is washed with hot water to dissolve only  $\text{PbCl}_2$ , confirmed further by another precipitation with  $\text{CrO}_4^{2-}$ .

Ion	Test Reagent or Method	Observation
$\text{Fe}^{3+}$	$\text{K}_3\text{Fe}(\text{CN})_6$	Dark blue precipitate
$\text{Cu}^{2+}$	$\text{NH}_3$	Dark blue solution
$\text{Cl}^-$	Add $\text{Ag}^+$	White precipitate ( $\text{AgCl}$ )
$\text{CO}_3^{2-}$	Add acid	Nonflammable gas evolved ( $\text{CO}_2$ )
$\text{SO}_4^{2-}$	Add $\text{Ba}^{2+}$	White precipitate ( $\text{BaSO}_4$ )

## ① SOLUBILITY RULES

Really, any list online works. I have found that this one is one of the most comprehensive.

①  $\text{AgF}$  is soluble.  $\text{PbCl}_2/\text{PbBr}_2/\text{PbI}_2$  mildly soluble at high temp.  
 $\text{Ag}_2\text{SO}_4$  is slightly soluble.

②  $\text{Li}_2\text{CO}_3$ ,  $\text{Li}_3\text{PO}_4$ ,  $\text{LiF}$  are insoluble.

③ Equally matched charges of 2 or greater are insoluble  
(eg.  $\text{PbSO}_4$ ,  $\text{CaSO}_4$ ,  $\text{BaSO}_4$ ,  $\text{AlPO}_4$ ,  $\text{BaCO}_3$ , etc. are insoluble)

④ Compounds that contain

•  $\text{NH}_4^+$ ,  $\text{Li}^+$ ,  $\text{Na}^+$ ,  $\text{K}^+$ ,  $\text{Rb}^+$ ,  $\text{Cs}^+$

•  $\text{NO}_3^-$ ,  $\text{C}_2\text{H}_3\text{O}_2^-$  ( $\text{CH}_3\text{COO}^-$ ),  $\text{SbF}_6^-$ ,  $\text{PF}_6^-$ ,  $\text{SiF}_6^{2-}$ ,  $\text{ClO}_3^-$ ,  $\text{ClO}_4^-$

any of these ions are soluble.

⑤  $\text{Pb}^{2+}$ ,  $\text{Ag}^+$ ,  $\text{Hg}^{2+}$ ,  $\text{Hg}_2^{2+}$  are insoluble.

⑥  $\text{F}^-$ ,  $\text{OH}^-$ ,  $\text{S}^{2-}$ ,  $\text{CO}_3^{2-}$ ,  $\text{PO}_4^{3-}$  are insoluble.

⑦  $\text{SO}_4^{2-}$ ,  $\text{Cl}^-$ ,  $\text{Br}^-$ ,  $\text{I}^-$  is soluble.

\* How to use this list: Go down this list in order until you find a rule that matches. You MUST apply these rules in order!!

## ② FLAME TESTS

Certain Cations in solution appear as a certain color when excited by fire energetically:

-  $\text{Ca}^{2+}$ ,  $\text{Li}^+$ ,  $\text{Sr}^{2+}$  → red

-  $\text{Fe}^{2+,3+}$ ,  $\text{Na}^+$  → yellow (Na impurities usually drown out all other colors in a flame test)

-  $\text{B}^{3+}$ ,  $\text{Ba}^{2+}$ ,  $\text{Cu}^{+,2+}$  → green

-  $\text{K}^+$ ,  $\text{Rb}^+$ ,  $\text{Cs}^+$  → blue-purple

-  $\text{Mg}^{2+}$  → very white