

# Gallium

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**Gallium** is a chemical element with symbol **Ga** and atomic number 31. It is in group 13 of the periodic table, and thus has similarities to the other metals of the group, aluminium, indium, and thallium. Gallium does not occur as a free element in nature, but as gallium(III) compounds in trace amounts in zinc ores and in bauxite. Elemental gallium is a soft, silvery blue metal at standard temperature and pressure, a brittle solid at low temperatures, and a liquid at temperatures greater than 29.76 °C (85.57 °F) (slightly above room temperature). The melting point of gallium is used as a temperature reference point. The alloy galinstan (68.5% gallium, 21.5% indium, and 10% tin) has an even lower melting point of −19 °C (−2 °F), well below the freezing point of water.

Since its discovery in 1875, gallium has been used to make alloys with low melting points. It is also used in semiconductors as a dopant in semiconductor substrates.

Gallium is predominantly used in electronics. Gallium arsenide, the primary chemical compound of gallium in electronics, is used in microwave circuits, high-speed switching circuits, and infrared circuits. Semiconductive gallium nitride and indium gallium nitride produce blue and violet light-emitting diodes (LEDs) and diode lasers. Gallium is also used in the production of artificial gadolinium gallium garnet for jewelry.

Gallium has no known natural role in biology. Gallium(III) behaves in a similar manner to ferric salts in biological systems, and has been used in some medical applications, including pharmaceuticals and radiopharmaceuticals. Gallium is used in thermometers as a non-toxic and environmentally friendly alternative to mercury and can withstand higher temperatures than mercury.

## Physical properties

Elemental gallium is not found in nature, but it is easily obtained by smelting. Very pure gallium metal has a silvery color and its solid metal fractures conchoidally like glass. Gallium liquid expands by 3.1% when it solidifies; therefore, it should not be stored in glass or metal containers because the container may rupture when the

### Gallium, <sup>31</sup>Ga



#### General properties

<b>Name, symbol</b>	gallium, Ga
<b>Appearance</b>	silvery blue

#### Gallium in the periodic table

<b>Atomic number</b> ( <i>Z</i> )	31
<b>Group, block</b>	group 13, p-block
<b>Period</b>	period 4
<b>Element category</b>	<span>□</span> post-transition metal
<b>Standard atomic weight</b> ( <i>A</i> <sub>r</sub> )	69.723(1) <sup>[1]</sup>
<b>Electron configuration</b>	[Ar] 3d <sup>10</sup> 4s <sup>2</sup> 4p <sup>1</sup>
<b>per shell</b>	2, 8, 18, 3

#### Physical properties

<b>Phase</b>	solid
<b>Melting point</b>	302.9146 K (29.7646 °C, 85.5763 °F)
<b>Boiling point</b>	2673 K (2400 °C,



Crystallization of gallium from the melt

gallium changes state. Gallium shares the higher-density liquid state with a short list of other materials that includes water, silicon, germanium, antimony, bismuth, and plutonium.<sup>[4]</sup>

Gallium attacks most other metals by diffusing into the metal lattice. For example, it diffuses into the grain boundaries of aluminium-zinc alloys<sup>[5]</sup> and steel,<sup>[6]</sup> making them very brittle. Gallium easily alloys with many metals, and is used in small quantities in the plutonium-gallium alloy in the plutonium cores of nuclear bombs to stabilize the plutonium crystal structure.<sup>[7]</sup>

The melting point of gallium, at 302.9146 K (29.7646 °C, 85.5763 °F), is just above room temperature, and is approximately the same as the average summer daytime temperatures in Earth's mid-latitudes. This melting point (mp) is one of the formal temperature reference points in the International Temperature Scale of 1990 (ITS-90) established by the International Bureau of Weights and Measures (BIPM).<sup>[8][9][10]</sup> The triple point of gallium, 302.9166 K (29.7666 °C, 85.5799 °F), is used by the US National Institute of Standards and Technology (NIST) in preference to the melting point.<sup>[11]</sup>

The unique melting point of gallium allows it to melt in the human hand, and then refreeze if removed. The liquid metal has a strong tendency to supercool below its melting point/freezing point. Seeding with a crystal helps to initiate freezing. Gallium is one of the four metals (with caesium, rubidium, and mercury) that are known to be liquid at, or near, normal room temperature. Of the four, gallium is the only one that is neither highly reactive nor highly toxic and can therefore be used in metal-in-glass high-temperature thermometers. It is also notable for having one of the largest liquid ranges for a metal, and for having (unlike mercury) a low vapor pressure at high temperatures. Gallium's boiling point, 2673 K, is more than eight times higher than its melting point on the absolute scale, the greatest ratio between melting point and boiling point of any element.<sup>[12]</sup> Unlike mercury, liquid gallium metal wets glass and skin, along with most other materials (with the exceptions of quartz, graphite, and Teflon), making it mechanically more difficult to

4352 °F)<sup>[2]</sup>

**Density** near r.t. 5.91 g/cm<sup>3</sup>

when liquid, at m.p. 6.095 g/cm<sup>3</sup>

**Heat of fusion** 5.59 kJ/mol

**Heat of vaporization** 256 kJ/mol<sup>[2]</sup>

**Molar heat capacity** 25.86 J/(mol·K)

#### Vapor pressure

P (Pa)	1	10	100	1 k	10 k	100 k
at T (K)	1310	1448	1620	1838	2125	2518

#### Atomic properties

**Oxidation states** 3, 2, 1, −1, −2, −4, −5<sup>[3]</sup>  
(an amphoteric oxide)

**Electronegativity** Pauling scale: 1.81

**Ionization energies**  
1st: 578.8 kJ/mol  
2nd: 1979.3 kJ/mol  
3rd: 2963 kJ/mol  
(more)

**Atomic radius** empirical: 135 pm

**Covalent radius** 122±3 pm

**Van der Waals radius** 187 pm

#### Miscellanea

**Crystal structure** orthorhombic



**Speed of sound** 2740 m/s (at 20 °C)  
thin rod

**Thermal expansion** 18 μm/(m·K) (at 25 °C)

handle even though it is substantially less toxic and requires far fewer precautions. Gallium painted onto glass is a brilliant mirror.<sup>[13]</sup> For this reason as well as the metal contamination and freezing-expansion problems, samples of gallium metal are usually supplied in polyethylene packets within other containers.

Gallium does not crystallize in any of the simple crystal structures. The stable phase under normal conditions is orthorhombic with 8 atoms in the conventional unit cell. Within a unit cell, each atom has only one nearest neighbor (at a distance of 244 pm). The remaining six unit cell neighbors are spaced 27, 30 and 39 pm farther away, and they are grouped in pairs with the same distance.<sup>[15]</sup> Many stable and metastable phases are found as function of temperature and pressure.<sup>[16]</sup>

The bonding between the two nearest neighbors is covalent; hence Ga<sub>2</sub> dimers are seen as the fundamental building blocks of the crystal. This explains the low melting point relative to the neighbor elements, aluminium and indium. This structure is strikingly similar to that of iodine and forms because of interactions between the single 4p electrons of gallium atoms, further away from the nucleus than the 4s electrons and the [Ar]3d<sup>10</sup> core. This phenomenon recurs with mercury with its "pseudo-noble-gas" [Xe]4f<sup>14</sup>5d<sup>10</sup>6s<sup>2</sup> electron configuration, which is liquid at room temperature.<sup>[17]</sup> The 3d<sup>10</sup> electrons do not shield the outer electrons very well from the nucleus and hence the first ionisation energy of gallium is greater than that of aluminium.<sup>[4]</sup>

The physical properties of gallium are highly anisotropic, i.e. have different values along the three major crystallographical axes *a*, *b*, and *c* (see table), producing a significant difference between the linear ( $\alpha$ ) and volume thermal expansion coefficients. The properties of gallium are strongly temperature-dependent, particularly near the melting point. For example, the coefficient of thermal expansion increases by several hundred percent upon melting.<sup>[14]</sup>

### Isotopes

Gallium has two stable isotopes: gallium-69 and gallium-73. The most commercially important radioisotopes are gallium-67 and gallium-68.

<b>Thermal conductivity</b>	40.6 W/(m·K)				
<b>Electrical resistivity</b>	270 nΩ·m (at 20 °C)				
<b>Magnetic ordering</b>	diamagnetic				
<b>Young's modulus</b>	9.8 GPa				
<b>Poisson ratio</b>	0.47				
<b>Mohs hardness</b>	1.5				
<b>Brinell hardness</b>	56.8–68.7 MPa				
<b>CAS Number</b>	7440-55-3				
<b>History</b>					
<b>Naming</b>	after <i>Gallia</i> (Latin for: France), homeland of the discoverer				
<b>Prediction</b>	Dmitri Mendeleev (1871)				
<b>Discovery and first isolation</b>	Lecoq de Boisbaudran (1875)				
<b>Most stable isotopes of gallium</b>					
<b>iso</b>	<b>NA</b>	<b>half-life</b>	<b>DM</b>	<b>DE (MeV)</b>	<b>DP</b>
<b>66Ga</b>	syn	9.5 h	β+	4.153	66Zn
<b>67Ga</b>	syn	3.261 d	ε	1.001	67Zn
<b>68Ga</b>	syn	1.13 h	β+	1.899	68Zn
<b>69Ga</b>	60.11%	is stable with 38 neutrons			
<b>70Ga</b>	syn	21.1 min	β−	1.653	70Ge
			ε	0.655	70Zn
<b>71Ga</b>	39.89%	is stable with 40 neutrons			
<b>72Ga</b>	syn	14.10 h	β−	3.997	72Ge
<b>73Ga</b>	syn	4.9 h	β−	1.598	73Ge

## Chemical properties

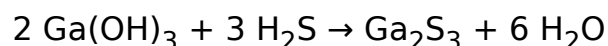
Gallium is found primarily in the +3 oxidation state. The +1 oxidation is also found in some compounds. For example, the very stable  $\text{GaCl}_2$  contains both gallium(I) and gallium(III) and can be formulated as  $\text{Ga}^{\text{I}}\text{Ga}^{\text{III}}\text{Cl}_4$ ; in contrast, the monochloride is unstable above 0 °C disproportionating into elemental gallium and gallium(III) chloride. Compounds containing gallium-gallium bonds are true gallium(II) compounds; for example, GaS can be formulated  $\text{Ga}_2^{4+}(\text{S}^{2-})_2$ , and the dioxan complex

$\text{Ga}_2\text{Cl}_4(\text{C}_4\text{H}_8\text{O}_2)_2$  contains a Ga–Ga bond.<sup>[18]</sup>

## Chalcogen compounds

Gallium reacts with the chalcogens only at relatively high temperatures. At room temperature, gallium metal is not reactive with air and water because it forms a passive, protective oxide layer. At higher temperatures, however, it reacts with atmospheric oxygen to form gallium(III) oxide,  $\text{Ga}_2\text{O}_3$ .<sup>[19]</sup> Reducing  $\text{Ga}_2\text{O}_3$  with elemental gallium in vacuum at 500 °C to 700 °C yields the dark brown gallium(I) oxide,  $\text{Ga}_2\text{O}$ .<sup>[20]:285</sup>  $\text{Ga}_2\text{O}$  is a very strong reducing agent, capable of reducing  $\text{H}_2\text{SO}_4$  to  $\text{H}_2\text{S}$ .<sup>[20]:207</sup> It disproportionates at 800 °C back to gallium and  $\text{Ga}_2\text{O}_3$ .<sup>[21]</sup>

Gallium(III) sulfide,  $\text{Ga}_2\text{S}_3$ , has 3 possible crystal modifications.<sup>[21]:104</sup> It can be made by the reaction of gallium with hydrogen sulfide ( $\text{H}_2\text{S}$ ) at 950 °C.<sup>[20]:162</sup> Alternatively,  $\text{Ga}(\text{OH})_3$  can be used at 747 °C.<sup>[22]</sup>



Reacting a mixture of alkali metal carbonates and  $\text{Ga}_2\text{O}_3$  with  $\text{H}_2\text{S}$  leads to the formation of *thiogallates* containing the  $[\text{Ga}_2\text{S}_4]^{2-}$  anion. Strong acids decompose these salts, releasing  $\text{H}_2\text{S}$  in the process.<sup>[21]:104–105</sup> The mercury salt,  $\text{HgGa}_2\text{S}_4$ , can be used as a phosphor.<sup>[23]</sup>

Gallium also forms sulfides in lower oxidation states, such as gallium(II) sulfide and the green gallium(I) sulfide, the latter of which is produced from the former by heating to 1000 °C under a stream of nitrogen.<sup>[21]:94</sup>

The other binary chalcogenides,  $\text{Ga}_2\text{Se}_3$  and  $\text{Ga}_2\text{Te}_3$ , have zincblende structure. They are all semiconductors but are easily hydrolysed and have limited utility.<sup>[21]:104</sup>

## Aqueous chemistry

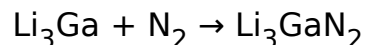
Strong acids dissolve gallium, forming gallium(III) salts such as  $\text{Ga}_2(\text{SO}_4)_3$  (gallium sulfate) and  $\text{Ga}(\text{NO}_3)_3$  (gallium nitrate). Aqueous solutions of gallium(III) salts contain the hydrated gallium ion,  $[\text{Ga}(\text{H}_2\text{O})_6]^{3+}$ .<sup>[24]:1033</sup> Gallium(III) hydroxide,  $\text{Ga}(\text{OH})_3$ , may be precipitated from gallium(III) solutions by adding ammonia. Dehydrating  $\text{Ga}(\text{OH})_3$  at 100 °C produces gallium oxide hydroxide,  $\text{GaO}(\text{OH})$ .<sup>[20]:140–141</sup>

Alkaline hydroxide solutions dissolve gallium, forming *gallate* salts containing the  $\text{Ga}(\text{OH})_4^-$  anion.<sup>[19][24]:1033[25]</sup> Gallium hydroxide, which is amphoteric, also dissolves in alkali to form gallate salts.<sup>[20]:141</sup> Although earlier work suggested  $\text{Ga}(\text{OH})_6^{3-}$  was another possible gallate anion,<sup>[26]</sup> it was not found in later work.<sup>[25]</sup>

## Pnictogen compounds

Gallium reacts with ammonia at 1050 °C to form gallium nitride,  $\text{GaN}$ . Gallium also forms binary compounds with phosphorus, arsenic, and antimony: gallium phosphide ( $\text{GaP}$ ), gallium arsenide ( $\text{GaAs}$ ), and gallium antimonide ( $\text{GaSb}$ ). These compounds have the same structure as  $\text{ZnS}$ , and have important semiconducting properties.<sup>[24]:1034</sup>  $\text{GaP}$ ,  $\text{GaAs}$ , and  $\text{GaSb}$  can be synthesized by the direct reaction of gallium with elemental phosphorus, arsenic, or antimony.<sup>[21]:99</sup> They exhibit higher electrical conductivity than  $\text{GaN}$ .<sup>[21]:101</sup>  $\text{GaP}$  can also be synthesized by reacting  $\text{Ga}_2\text{O}$  with phosphorus at low temperatures.<sup>[27]</sup>

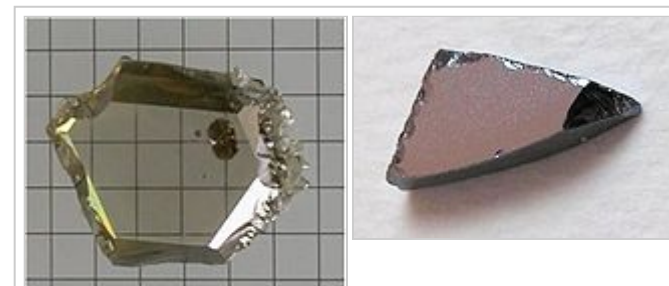
Gallium forms ternary nitrides; for example:<sup>[21]:99</sup>



Similar compounds with phosphorus and arsenic are possible:  $\text{Li}_3\text{GaP}_2$  and  $\text{Li}_3\text{GaAs}_2$ . These compounds are easily hydrolyzed by dilute acids and water.<sup>[21]:101</sup>

## Halides

Gallium(III) oxide reacts with fluorinating agents such as  $\text{HF}$  or  $\text{F}_2$  to form gallium(III) fluoride,  $\text{GaF}_3$ . It is an ionic compound strongly insoluble in water. However, it dissolves in hydrofluoric acid, in which it forms an adduct with water,  $\text{GaF}_3 \cdot 3\text{H}_2\text{O}$ . Attempting to dehydrate this adduct forms  $\text{GaF}_2\text{OH} \cdot n\text{H}_2\text{O}$ . The adduct reacts with ammonia to form  $\text{GaF}_3 \cdot 3\text{NH}_3$ , which can then be heated to form anhydrous  $\text{GaF}_3$ .<sup>[20]:128–129</sup>

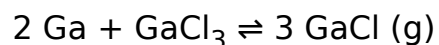


Gallium nitride (left) and gallium arsenide (right) crystals

Gallium trichloride is formed by the reaction of gallium metal with chlorine gas.<sup>[19]</sup> Unlike the trifluoride, gallium(III) chloride exists as dimeric molecules, Ga<sub>2</sub>Cl<sub>6</sub>, with a melting point of 78 °C. Equivalent compounds are formed with bromine and iodine, Ga<sub>2</sub>Br<sub>6</sub> and Ga<sub>2</sub>I<sub>6</sub>.<sup>[20]:133</sup>

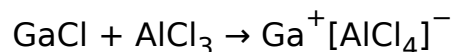
Like the other group 13 trihalides, gallium(III) halides are Lewis acids, reacting as halide acceptors with alkali metal halides to form salts containing GaX<sub>4</sub><sup>−</sup> anions, where X is a halogen. They also react with alkyl halides to form carbocations and GaX<sub>4</sub><sup>−</sup>.<sup>[20]:136–137</sup>

When heated to a high temperature, gallium(III) halides react with elemental gallium to form the respective gallium(I) halides. For example, GaCl<sub>3</sub> reacts with Ga to form GaCl:

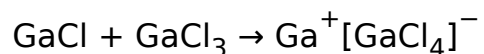


At lower temperatures, the equilibrium shifts toward the left and GaCl disproportionates back to elemental gallium and GaCl<sub>3</sub>. GaCl can also be produced by reacting Ga with HCl at 950 °C; the product can be condensed as a red solid.<sup>[24]:1036</sup>

Gallium(I) compounds can be stabilized by forming adducts with Lewis acids. For example:



The so-called "gallium(II) halides", GaX<sub>2</sub>, are actually adducts of gallium(I) halides with the respective gallium(III) halides, having the structure Ga<sup>+</sup>[GaX<sub>4</sub>]<sup>−</sup>. For example:<sup>[19][24]:1036[28]</sup>



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