

Water Chemistry

Matthew Williams • Chemistry • May 15, 2026

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Water is the most important substance in chemistry and in life. Its unusual physical properties all arise from a single structural feature: the ability to form hydrogen bonds. Understanding what makes water hard, how hardness is removed, and how water is made safe to drink are core examination topics.

Unique Properties of Water

Water's unusual properties arise from hydrogen bonds between molecules — each O–H group in one molecule attracts the lone pair on the oxygen of a neighbouring molecule. These bonds are weak individually but collectively make water unlike almost any other molecular compound:

Property	Significance
High specific heat capacity	Absorbs a lot of heat for a small temperature rise — moderates climate and body temperature
High boiling point (100 °C)	Remains liquid over a wide temperature range suitable for life
Excellent solvent	Dissolves ionic and polar substances, allowing nutrients and waste to be transported in living systems
Density maximum at 4 °C	Ice is less dense than liquid water — ice floats, insulating aquatic organisms beneath
High surface tension	Allows surface-dwelling insects; important in capillary action in plants

All these properties would be very different if water molecules could not form hydrogen bonds. A molecule of similar size without H-bonding (e.g. H_2S) is a gas at room temperature and boils at $-60\text{ }^\circ\text{C}$.

Hardness of Water

Not all water behaves the same way. Water that flows through limestone or chalk picks up dissolved calcium and magnesium ions, and these affect how it reacts with soap and what it deposits in pipes and kettles.

Hard water does not lather readily with soap. It contains dissolved Ca^{2+} and Mg^{2+} ions, usually picked up as rainwater passes through limestone or dolomite rock. When soap (a sodium salt of a fatty acid) is added, these ions react with it to form an insoluble scum instead of a lather.

Temporary Hardness

Cause: dissolved calcium hydrogen carbonate, $\text{Ca}(\text{HCO}_3)_2$. This forms when rainwater containing dissolved CO_2 passes through limestone:

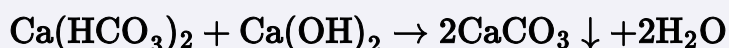


Temporary hardness is removed by:

- **Boiling** — reverses the above reaction, precipitating CaCO_3 (this is what forms limescale in kettles and pipes):



- **Adding lime, $\text{Ca}(\text{OH})_2$** , — also precipitates CaCO_3

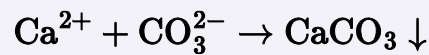


Permanent Hardness

Cause: dissolved calcium or magnesium sulfates (CaSO_4 , MgSO_4). These do not decompose on boiling, so permanent hardness cannot be removed by boiling.

Permanent (and temporary) hardness is removed by:

- **Adding washing soda (Na_2CO_3)** — the carbonate ion precipitates Ca^{2+} and Mg^{2+} as their insoluble carbonates:



- **Ion exchange resins** — the hard water passes through a resin that exchanges Ca^{2+} and Mg^{2+} for Na^+ or H^+ ions, which do not cause hardness
- **Distillation** — removes all dissolved solids, giving pure water

Remember

Boiling only removes **temporary** hardness — it has no effect on permanent hardness. Washing soda and ion exchange remove both. This distinction appears frequently on Paper 01.

Advantages and Disadvantages of Hard Water

Hard water is not entirely undesirable. Ca^{2+} ions contribute to strong bones and teeth, and hard water reduces the leaching of lead from old pipes (a thin calcium carbonate layer forms on the inside). The main disadvantages are: soap is wasted forming scum, limescale builds up in heating systems reducing efficiency, and hot water systems corrode more quickly.

Leaching

Water's ability to dissolve ionic and polar substances has a consequence in soil: **leaching** is the process by which dissolved minerals and nutrients are washed out of the topsoil and carried downward by rainwater into deeper layers or into groundwater. Nitrates, potassium ions, and other soluble nutrients can leach out of agricultural soils, particularly after heavy rainfall. This:

- Depletes soil fertility, reducing crop yields
- Carries nitrates and phosphates into rivers and lakes, causing eutrophication
- Contaminates groundwater used for drinking

Leaching is why farmers on sandy or well-drained soils need to apply fertilisers more often — the nutrients do not stay in the root zone as long as they would in clay-heavy soils, which hold water (and dissolved ions) more strongly.

Water Treatment for Drinking

Softening hard water solves a domestic inconvenience. Making water safe to drink is a public health necessity, and the two problems are handled differently.

Natural water sources contain suspended solids, microorganisms, and dissolved impurities. Treatment makes it safe:

Stage	Process	Purpose
Sedimentation	Allow particles to settle (sometimes with aluminium sulfate as a flocculant to clump fine particles)	Removes coarse suspended matter
Filtration	Pass through sand or gravel beds	Removes remaining suspended particles and some microorganisms
Chlorination	Add chlorine gas or sodium hypochlorite	Kills bacteria and other pathogens
pH adjustment	Add lime (Ca(OH)) if too acidic	Prevents corrosion of pipes

Boiling water kills pathogens and is the most accessible treatment in an emergency. It does not remove dissolved salts, heavy metals, or chemical pollutants.

Exam Tip

Be precise about the order of treatment stages: sedimentation before filtration (you remove the large particles first, then the fine ones). Chlorination comes after filtration.