

Topic 1.8: Valence Electrons and Ionic Compounds

Revision Notes · AP Chemistry

UNIT 1 · ATOMIC STRUCTURE AND PROPERTIES

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Why Atoms Form Ions At All

An atom's chemistry is run by its valence electrons, and the most stable arrangement an atom can reach is a full outer shell, the configuration of a noble gas.

Atoms get there by the shortest route. A metal with one, two or three valence electrons loses them all: the shell beneath is already full, so the ion is born stable. A nonmetal that is one, two or three electrons short gains exactly that many, completing its outer shell.

That is why ionic charges are predictable from the periodic table alone. The charge is not arbitrary; it is the arithmetic of reaching the nearest full shell.

The Charge Pattern Across the Table

Groups 1 and 2. Group 1 atoms have one valence electron and form 1+ ions: Na^+ , K^+ . Group 2 atoms have two and form 2+ ions: Mg^{2+} , Ca^{2+} , Ba^{2+} . Aluminium, in Group 13, loses three: Al^{3+} .

Groups 16 and 17. Group 17 atoms are one electron short and form 1– ions: F^- , Cl^- , Br^- . Group 16 atoms are two short and form 2– ions: O^{2-} , S^{2-} . Nitrogen and phosphorus, Group 15, form 3– ions: N^{3-} , P^{3-} .

The middle stays variable. Transition metals lose their ns electrons first and then a variable number of (n–1)d electrons, so many form more than one stable cation: Fe^{2+} and Fe^{3+} , Cu^+ and Cu^{2+} , Pb^{2+} and Pb^{4+} . Silver (Ag^+) and zinc (Zn^{2+}) are the reliable exceptions with one fixed charge each.

Names signal the charge. Cations keep the element's name (sodium ion, iron(III) ion, where the Roman numeral states the charge for variable metals). Simple anions change the ending to -ide: chloride, oxide, nitride, sulfide.

Worked Example: From Configuration to Ion

Magnesium, $Z = 12$, is $[\text{Ne}] 3s^2$. Losing its two 3s electrons leaves $[\text{Ne}]$, the stable neon core, and the ion is Mg^{2+} because two electrons (negative) are gone and the 12 protons remain.

Sulfur, $Z = 16$, is $[\text{Ne}] 3s^2 3p^4$. Gaining two electrons fills the 3p subshell: $[\text{Ne}] 3s^2 3p^6$, the argon configuration, and the ion is S^{2-} .

In each case the ion is isoelectronic with a noble gas, which is the whole point. The metal empties its outer shell; the nonmetal fills its own; both end at a full shell by opposite routes.

Building the Formula of an Ionic Compound

An ionic compound is neutral overall, so the total positive charge must exactly balance the total negative charge. That single condition writes the formula.

The method. Find the smallest numbers of each ion that make the charges balance, which is the lowest common multiple of the two charge sizes. Those numbers become the subscripts, reduced to the lowest ratio.

Worked examples. Calcium ($2+$) with chlorine ($1-$): one Ca^{2+} balances two Cl^- , giving CaCl_2 . Aluminium ($3+$) with oxygen ($2-$): the lowest common multiple of 3 and 2 is 6, so two Al^{3+} ($+6$) balance three O^{2-} (-6), giving Al_2O_3 . Magnesium ($2+$) with nitrogen ($3-$): three Mg^{2+} ($+6$) balance two N^{3-} (-6), giving Mg_3N_2 .

Charges cancel downward too. Calcium ($2+$) with oxygen ($2-$) gives CaO , not Ca_2O_2 . Subscripts in an ionic formula are always the lowest whole number ratio, exactly as an empirical formula is.

Reading formulas backwards. The formula tells you the ions. FeCl_2 must contain Fe^{2+} , because two chloride ions (-1 each) need $+2$ to balance; FeCl_3 contains Fe^{3+} . For variable charge metals, the anion count is how you recover the cation's charge, and the Roman numeral in the name reports it.

Traps That Cost Marks in Topic 1.8

Giving transition metals a single fixed charge. Iron is not simply $3+$ or $2+$; both ions are common, and a question's formula or name decides which one is present. Check the anion total before naming the charge.

Writing Ca_2O_2 or similar unreduced formulas. Ionic formulas are lowest ratios. If both subscripts share a factor, divide it out. A $2+$ ion with a $2-$ ion always gives a 1:1 formula.

Balancing atoms instead of charges. The rule is charge neutrality, not equal numbers of ions. CaCl_2 has twice as many chloride ions as calcium ions, and that is precisely what neutrality requires.

Losing d electrons first in transition metal ions. As in Topic 1.5, the ns electrons leave before the $(n-1)d$ electrons. Fe^{2+} is $[\text{Ar}] 3d^6$; an explanation that removes 3d electrons first contradicts the configuration evidence.

Predicting charges for elements that do not form simple ions. Carbon and silicon (Group 14) would need to gain or lose four electrons; the energy cost is prohibitive, so they share electrons in covalent compounds instead. Simple ionic charges run from $3-$ to $3+$ for main group elements, and hydrogen is special: $+1$ with nonmetals, and -1 (hydride) with active metals.

Confusing the ion's electron count. Mg^{2+} has 10 electrons, not 12 or 14. Protons never change when an ion forms; only the electron count moves, up for anions and down for cations.

Key Relationships to Remember

- Main group ionic charges: Group 1 $\rightarrow 1+$, Group 2 $\rightarrow 2+$, Al $\rightarrow 3+$, Group 15 $\rightarrow 3-$, Group 16 $\rightarrow 2-$, Group 17 $\rightarrow 1-$; Ag $\rightarrow 1+$, Zn $\rightarrow 2+$
- Atoms form the ion that reaches the nearest full outer shell; the resulting ion is isoelectronic with a noble gas
- Ionic formula: subscripts are the smallest numbers that make total positive charge equal total negative charge, reduced to the lowest ratio
- Variable charge metals ($\text{Fe}^{2+}/\text{Fe}^{3+}$, $\text{Cu}^+/\text{Cu}^{2+}$) state their charge with a Roman numeral; recover it from the anion count in the formula
- Ion formation changes only the electron count; the proton count, and so the element, never changes

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