

Unit

1

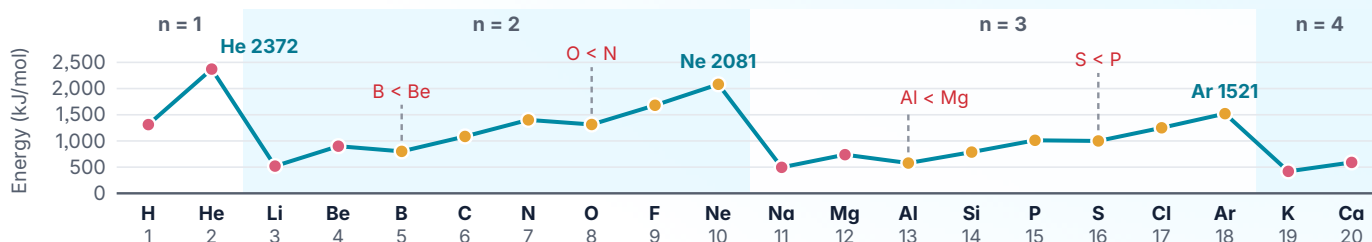
AP Chemistry

Atomic Structure and Properties

Moles · Mass spectra · Composition · Mixtures · Electron configuration · Photoelectron spectroscopy · Periodic trends · Ionic compounds

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basesUnit 9
7–9%
Thermo-
dynamics,
electro-
chemistry

Unit 1 is 7–9% of the multiple-choice section. No free-response question is only about Unit 1. Most years one part asks you to explain a difference in atomic size or ionization energy, and the mole calculations from this unit are a step in many other questions.

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First ionization energy drops sharply at Li, Na and K because each starts a new shell, farther from the nucleus. It rises across a period as more protons pull on the same shell. The four marked dips have their own reasons: the outer electron of B and Al is the first in a higher-energy p subshell, and the outer electron of O and S is the first to pair up in a p orbital.

● s block (groups 1 and 2, and He) ● p block (groups 13 to 18) n = the shell that holds the outer electrons

What the exam asks most

- Explain why one atom or ion is bigger, or easier to ionize, than another.** Name the shell (n) of each one's outer electrons: a higher shell is farther out and attracted less strongly. In the same shell, more protons pull harder. "It's the periodic trend" earns no credit. (1.7)
- Write the electron configuration of an atom or an ion.** Write each subshell with its electron count as a superscript. Row-4 transition metals lose their 4s electrons first: Fe is $[\text{Ar}] 4s^2 3d^6$, so Fe^{2+} is $[\text{Ar}] 3d^6$. (1.5)
- Find how pure a sample is.** Change the precipitate's mass to moles, use the formula to get moles of the substance, then grams, and divide by the sample's mass. For a lab error, say how it changes the mass weighed, then the percent. (1.4)

- Turn masses into moles and formulas.** Build each molar mass from the periodic table. Use the formula as a ratio: 1 mol of H_2O holds 2 mol of H. For an empirical formula, divide every mole amount by the smallest. (1.1) (1.3)
- Read a photoelectron spectrum.** Each peak is a subshell; higher-energy peaks hold electrons closer to the nucleus. A peak's relative height gives its number of electrons, and in a neutral atom the heights add up to the atomic number. (1.6)
- Read or draw a mass spectrum.** Each peak is one isotope, at its mass and as tall as its percent. The percents add up to 100, and neutrons = mass number – atomic number. (1.2)

Not on the exam

You won't need an electron's four quantum numbers (only the shell number n), exceptions like Cr and Cu, or mass spectra of molecules or element mixtures.

Formulas and facts to know

Given on the equation sheet: $E = h\nu$ and $c = \lambda\nu$, where ν (nu) is the frequency, λ (lambda) the wavelength, h Planck's constant and c the speed of light; $n = m/M$ (moles = mass ÷ molar mass); Avogadro's number, $6.022 \times 10^{23} \text{ mol}^{-1}$; Coulomb's law, $F \propto q_1q_2/r^2$ (the force grows with the charges q and weakens with the distance r); and a periodic table with each element's symbol, atomic number and average atomic mass, but no names, ion charges or radii.

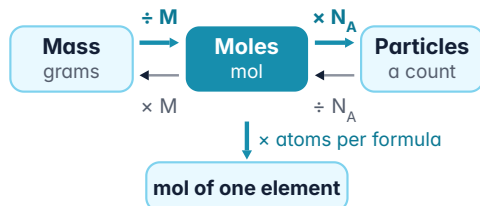
Not given, so learn these:

- The filling order: 1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p (4s fills before 3d).
- Ions of the transition metals in row 4 lose their 4s electrons before any 3d electrons.
- A cation is smaller than its atom; an anion is larger.
- The usual ion charges: group 1 forms +1 ions, group 2 forms +2, group 13 +3, group 15 –3, group 16 –2 and group 17 –1.
- An atom's mass in amu (atomic mass units) is the same number as the mass of a mole of those atoms in grams: 35.45 amu per Cl atom, 35.45 g per mole.
- Questions give you any ionization energies, photoelectron energies and radii they want you to use.

1.1 Moles and molar mass

You can't count atoms at the bench, so you weigh them. The mole turns a mass you can measure into a number of particles you can't see.

M = molar mass (g/mol), from the periodic table



What a mole is

1 mol = 6.022×10^{23} particles: atoms, molecules or formula units. This is Avogadro's number, N_A .

The average mass of one particle in amu (atomic mass units) is the same number as the molar mass in g/mol. One Cl atom averages **35.45 amu**, so a mole of Cl atoms weighs **35.45 g**. For moles of one element in a compound, multiply by its subscript: 1 mol of AlCl_3 holds 3 mol Cl.

On the equation sheet

$$n = \frac{m}{M} \quad \text{moles} = \text{mass} \div \text{molar mass}$$

Worked example

How many grams of chlorine are in 1.25 mol of AlCl_3 ?

- Each AlCl_3 holds 3 Cl, so use the formula as a ratio:

$$1.25 \text{ mol AlCl}_3 \times \frac{3 \text{ mol Cl}}{1 \text{ mol AlCl}_3} = 3.75 \text{ mol Cl}$$

- Then convert to mass: $3.75 \text{ mol} \times 35.45 \text{ g/mol} = 132.9 \text{ g}$

133 g Cl (three significant figures)

In 1 mol of $\text{Ca}(\text{NO}_3)_2$	mol
Ca^{2+} ions	1
NO_3^- ions	2
N atoms	2
O atoms	6

Check that the answer is a sensible size

A mole of a substance weighs grams to hundreds of grams, and one atom 10^{-24} to 10^{-22} g. An answer far outside these has a power-of-ten slip.

Exam trap

- Using the mass of one atom for a molecule. Iodine is I_2 , so its molar mass is $2 \times 126.90 = 253.80 \text{ g/mol}$. 1.021 g of I_2 is $1.021 \div 253.80 = 0.004023 \text{ mol}$. Dividing by 126.90 doubles the answer.
- Treating grams per liter as moles per liter. $0.0016 \text{ g/L of NH}_4\text{Cl}$ means 0.0016 g in each liter. That is $0.0016 \div 51.48 = 3.1 \times 10^{-5} \text{ mol}$ in 1.0 L, not 0.0016 mol .
- Leaving out the subscript. One mole of Ag_2S contains two moles of Ag atoms, so $0.50 \text{ mol of Ag}_2\text{S}$ holds 1.0 mol of Ag.

Build every molar mass from the periodic table

Write the sum out: $\text{CH}_3\text{NH}_2\text{Cl}$ is $12.01 + 6(1.008) + 14.01 + 35.45 = 67.52 \text{ g/mol}$. A wrong molar mass can cost a point, but later steps that use your number correctly can still earn their points. Use the table's values as printed (H is 1.008, Mg is 24.30) without rounding.

Significant figures and units

- Which digits count. 2.200 g has four significant figures. 0.00250 mol has three (the 2, 5 and last 0). To show 1500 with two, write 1.5×10^3 .
- Multiplying or dividing: the answer keeps as many significant figures as the measurement with the fewest.
- Adding or subtracting: the answer keeps as many decimal places as the measurement with the fewest: $1.698 \text{ g} - 1.462 \text{ g} = 0.236 \text{ g}$.
- Exact numbers don't limit significant figures. Counted atoms, mole ratios and the 2 in H_2O have unlimited significant figures.
- Round once, at the end, and carry extra digits through every step. Put a unit on every number, including the answer.
- Decimal places are not significant figures. A balance that reads to 0.01 g gives 0.47 g , which has only two significant figures.

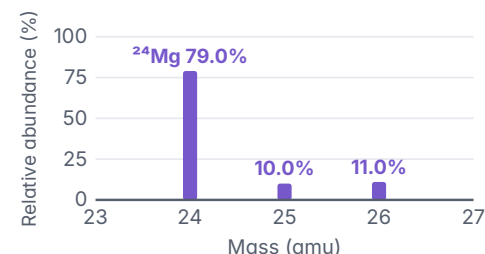
Quick check

Gallium has peaks at 69 amu (60.1%) and 71 amu (39.9%). Estimate its average atomic mass. How many neutrons are in the lighter isotope?

Answer: $0.601(69) + 0.399(71) = 69.8 \text{ amu}$ (the periodic table gives 69.72). Ga has 31 protons, so ^{69}Ga has $69 - 31 = 38$ neutrons.

1.2 Mass spectra of elements

A mass spectrometer sorts an element's atoms by mass. Each isotope makes one peak, placed at its mass, with a height that shows what percent of the atoms have that mass.



Magnesium. Every Mg atom has 12 protons and 12 electrons; the three isotopes have 12, 13 and 14 neutrons.

The top number is the mass number, protons plus neutrons. The bottom number is the atomic number, the number of protons. So this atom has $26 - 12 = 14$ neutrons.

Worked example

Chlorine is 75.76% ^{35}Cl (34.969 amu) and 24.24% ^{37}Cl (36.966 amu). Find its average atomic mass.

- Change each percent to a decimal and multiply by that isotope's mass:
 $0.7576(34.969) + 0.2424(36.966)$
- $= 26.49 + 8.961 = 35.45$

35.45 amu, closer to 35 than 37 because about three of every four Cl atoms are ^{35}Cl

Exam trap

- Thinking isotopes differ in protons or electrons. They differ only in neutrons. ^{25}Mg and ^{26}Mg both have 12 protons and 12 electrons; ^{26}Mg has 14 neutrons and ^{25}Mg has 13.
- Giving one number when asked for protons and neutrons. A silicon peak at 28 amu means 14 protons (silicon's atomic number) and $28 - 14 = 14$ neutrons. Write both.

Drawing missing peaks

Say a question shows only the ^{24}Mg line (79%) and says the other two isotopes are about equally common. The percents add to 100, so each is $(100 - 79) \div 2 = 10.5\%$. Check the y-axis scale, then draw both lines 10.5% tall, at 25 and 26 amu.

1.3 Elemental composition of pure substances

A pure compound always has the same composition. Every sample of water is 11.19% hydrogen by mass. This is the **law of definite proportions**.

Some pure substances are molecules (H_2O , CO_2). Others are ions held in a fixed ratio, written as a **formula unit** (NaCl , MgO). The **empirical formula** is the lowest whole-number ratio of atoms; the molecular formula is a whole-number multiple of it.

Mass percent

$$\% X = \frac{\text{mass of X in 1 mol}}{\text{molar mass}} \times 100$$

$$\text{H}_2\text{O}: \frac{2(1.008)}{18.02} \times 100 = 11.19\% \text{ H.}$$

H ₂ O	H	O 88.81%	
CO ₂	C 27.29%	O 72.71%	
Caffeine	C 49.47%	N 28.86%	O

An O atom (16.00 amu) outweighs water's two H atoms (2.016 amu) eight to one.

H C N O

From mass percents to a formula

- 1 Assume 100 g, so 40.0% becomes 40.0 g.
- 2 Change grams to moles by dividing by each element's molar mass.
- 3 Divide every mole amount by the smallest one.
- 4 If a result isn't near a whole number (.5, .33, .25 ...), multiply all of them by the same whole number (table below).
- 5 For the molecular formula, divide the molar mass by the empirical-formula mass and multiply every subscript by the result.

Ratio ends in	Multiply by	Example
.5	2	1 : 1.5 $\rightarrow$ 2 : 3
.33 or .67	3	1 : 1.33 $\rightarrow$ 3 : 4
.25 or .75	4	1 : 1.25 $\rightarrow$ 4 : 5

Worked example

How many grams of iron are in 10.0 g of Fe_2O_3 ?

- 1 Fraction of the mass that is Fe: $\frac{2(55.85)}{159.70} = 0.6994$, or 69.94%

$$10.0 \text{ g} \times 0.6994 = 6.99 \text{ g of Fe}$$

Worked example

A white solid is 49.47% C, 5.19% H, 28.86% N and 16.48% O by mass, and its molar mass is 194 g/mol. Find its empirical and molecular formulas.

- 1 Moles in 100 g:
C: $49.47 \div 12.01 = 4.119$
H: $5.19 \div 1.008 = 5.149$
N: $28.86 \div 14.01 = 2.060$
O: $16.48 \div 16.00 = 1.030$
- 2 Divide each by the smallest, 1.030:
C 3.999, H 4.999, N 2.000, O 1.000.
That gives $\text{C}_4\text{H}_5\text{N}_2\text{O}$, which has a mass of 97.10 g/mol.
- 3 $194 \div 97.10 = 2.00$, so double every subscript.

Empirical $\text{C}_4\text{H}_5\text{N}_2\text{O}$, molecular $\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2$ (caffeine)

Formulas from combustion data

Burning a compound of C, H and O turns all its carbon into CO_2 and all its hydrogen into H_2O , so:

$$\begin{aligned} \text{mol C} &= \text{mol CO}_2 \\ \text{mol H} &= 2 \times \text{mol H}_2\text{O} \end{aligned}$$

A 7.045 g sample gives 0.2400 mol CO_2 and 0.1600 mol H_2O : 0.2400 mol C (2.882 g) and 0.3200 mol H (0.323 g). The rest is oxygen: $7.045 - 2.882 - 0.323 = 3.840$ g, or 0.2400 mol. Dividing by 0.2400 gives C : H : O = 1 : 1.333 : 1, and tripling gives $\text{C}_3\text{H}_4\text{O}_3$.

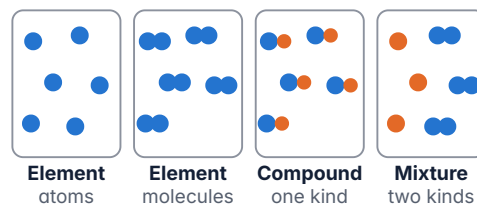
Exam trap

- Using moles of water as moles of hydrogen. Each H_2O has two H atoms, so 0.1600 mol of water holds 0.3200 mol of H.
- Rounding 1.5 up to 2. A ratio ending in .5 means multiply by 2. An iron oxide that is 69.94% Fe has 1.252 mol Fe and 1.879 mol O per 100 g, a ratio of 1 : 1.50. Doubling gives Fe_2O_3 ; rounding gives the wrong FeO_2 .
- Subtracting the wrong masses. If you weigh a beaker with manganese, let it react with chlorine and weigh again, the chlorine's mass is (beaker + Mn + Cl) - (beaker + Mn). Subtracting the empty beaker counts the Mn too.

Compounds that share an empirical formula

CH_2O is the empirical formula of formaldehyde (CH_2O), acetic acid ($\text{C}_2\text{H}_4\text{O}_2$) and glucose ($\text{C}_6\text{H}_{12}\text{O}_6$). Their mass percents are the same, so only the molar mass tells them apart.

1.4 Composition of mixtures



Each dot is an atom; touching dots form one particle. A **pure substance** has one kind of particle and a fixed composition. A **mixture** has two or more kinds, in varying proportions.

Finding the percent purity of a sample

- 1 Change the mass of the solid you weighed to moles.
- 2 Use the formula ratio to get moles of what you want (1 AgCl per NaCl).
- 3 Multiply by its molar mass to get grams.
- 4 Divide by the mass of the **original sample** and multiply by 100.

Worked example

A 1.000 g mixture of NaCl and sand gives 2.200 g of dry AgCl with excess AgNO_3 . What percent is NaCl?

- 1 $2.200 \div 143.32 = 0.01535$ mol of AgCl, which is also the moles of NaCl (one Cl each).
- 2 $0.01535 \times 58.44 = 0.8971$ g of NaCl

$$0.8971 \div 1.000 \times 100 = 89.71\% \text{ NaCl; the rest is sand}$$

Lab error	Mass weighed	Calculated result
Precipitate still wet	too high	too high
Precipitate splattered out	too low	too low
Filter paper not subtracted	too high	too high
Iron in a sample weighed as Fe_2O_3 , but some formed FeO	too low	% Fe too low

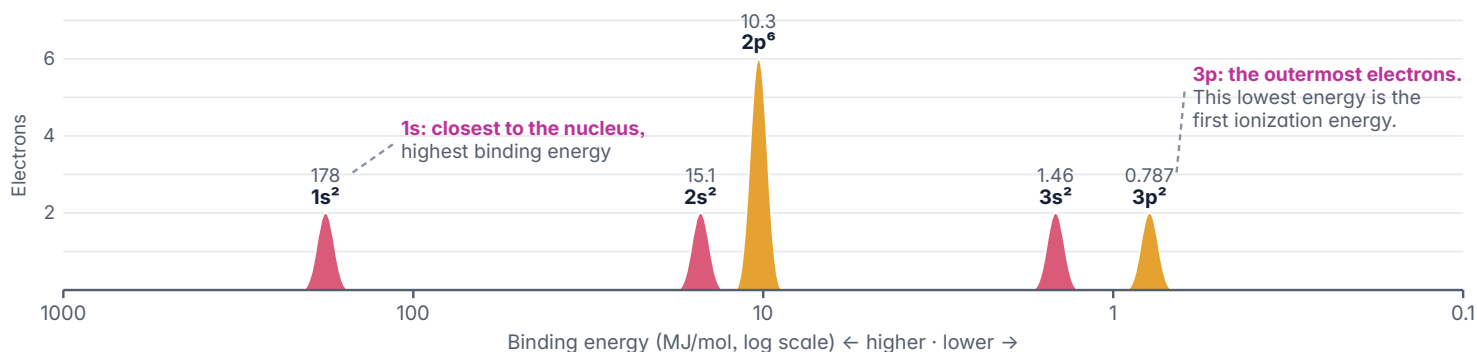
Exam trap

Saying "too low" without the reason. Give the chain: "Precipitate splattered out, so its mass was too low, so the moles and percent of NaCl came out too low." Say "no change" only if no measured mass changes.

To dry to **constant mass**, heat and reweigh until the mass stops changing; don't average the weighings.

1.6 Photoelectron spectroscopy (PES)

High-energy light knocks electrons out of atoms, and the energy each one needed shows which subshell it came from



Silicon, $1s^2 2s^2 2p^6 3s^2 3p^2$. Each peak is one subshell. A peak's position is the energy needed to remove an electron from that subshell (the binding energy, in MJ/mol; 1 MJ = 1,000 kJ). The axis is a log scale, so each labeled step is ten times the energy. A peak's height is the number of electrons in the subshell, and the heights add up to 14, silicon's atomic number.

■ s subshell ■ p subshell

How to read a spectrum

- 1 Check which way energy increases. Here it's to the left, so label the peaks from the left: 1s, 2s, 2p, 3s, 3p, 4s.
- 2 Turn heights into electrons, using the scale on the axis.
- 3 Add them. For a neutral atom the total is the atomic number.
- 4 Write the configuration and name the element.

Explaining where the 1s peak sits

The 1s electrons are **closest to the nucleus**, so they feel the **strongest Coulombic attraction** and need the most energy to remove.

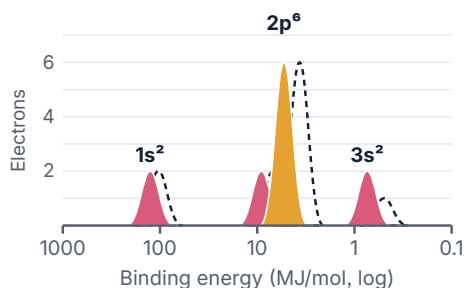
Exam trap

- Reading the energy axis as atomic number or mass. A peak near 40 MJ/mol doesn't mean zirconium (atomic number 40) or calcium (mass 40).
- Counting peaks as electrons. Each peak is a whole subshell: neon has three peaks but 10 electrons.
- Drawing every p peak 6 tall. Silicon has only two 3p electrons, so its 3p peak is 2 tall.

Energy per mole to energy per electron

Divide by Avogadro's number. Sodium's valence peak, 0.496 MJ/mol, is 4.96×10^5 J/mol, so one electron needs $4.96 \times 10^5 \div 6.022 \times 10^{23} = 8.24 \times 10^{-19}$ J. By $\lambda = hc/E$, light of 241 nm or shorter can remove it.

Comparing Na and Mg



The solid, labeled peaks are Mg (126, 9.07, 5.31 and 0.738 MJ/mol); the dashed outlines are Na (104, 6.84, 3.67 and 0.496).

Every Mg peak sits to the left of the matching Na peak, because 12 protons pull on each subshell harder than 11 do. Mg's 3s peak is twice as tall: Mg has two 3s electrons and Na has one. Ions work the same way: Na^+ has neon's peak heights (2, 2, 6), but its peaks sit left of neon's, because 11 protons pull on the same 10 electrons.

For the same subshell, more protons means a higher binding energy: the 1s electrons need 84.0 MJ/mol in neon (10 protons), 104 in sodium (11) and 390 in calcium (20). Within one atom, distance matters most: calcium's 4s electrons, far from the nucleus, need only 0.590 MJ/mol.

Drawing a missing peak

- Position. Put the outermost subshell's peak at the low-energy end. In an element with more protons, each core peak sits at a higher energy.
- Height. Match the electron count on the given scale. Silicon's 3p peak is 2 units tall, the same as its 3s.

On the equation sheet

$E = h\nu$ and $c = \lambda\nu$ (ν , nu, is the frequency and λ , lambda, the wavelength), so $\lambda = hc/E$

Worked example

Calcium's valence peak is at 9.80×10^{-19} J per electron. What wavelength of light can remove that electron?

- 1 Use the lowest-energy peak, the valence 4s electrons.
- 2 $\lambda = \frac{(6.626 \times 10^{-34})(2.998 \times 10^8)}{9.80 \times 10^{-19}}$

$$\lambda = 2.03 \times 10^{-7} \text{ m, or } 203 \text{ nm}$$

Exam trap

- Using the wrong peak. Removing the outermost electron takes the energy of the lowest-energy peak. Calcium's 1s peak gives a wavelength hundreds of times too short.
- Using the energy as if it were the frequency. Solve $E = h\nu$ for $\nu = E/h$ first, then use $\lambda = c/\nu$.
- Losing a power of ten. If you type " $\times 10^{\wedge}$ ", bracket the whole bottom of the fraction. For a valence electron, expect about 10^{-7} m.

Enter powers of ten

TI-84 Plus CE. Type 6.626, then $\boxed{2\text{nd}} \boxed{,}$ (EE, short for " $\times 10$ to the power"), then $\boxed{(-)} \boxed{34}$. The screen shows 6.626E-34. Use $\boxed{(-)}$ for a negative sign, never the subtraction key.

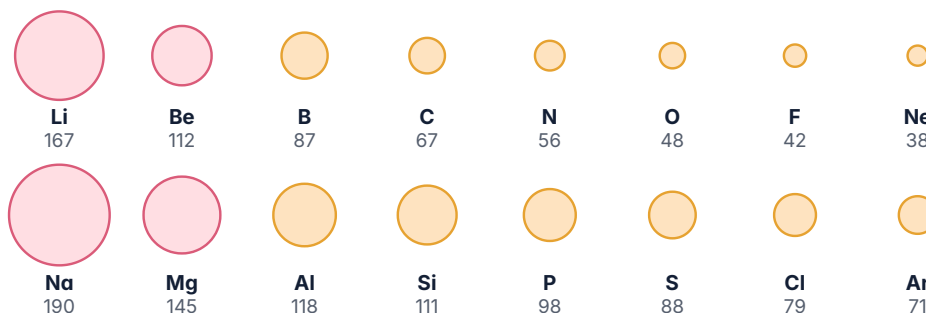
TI-Nspire CX. Type 6.626 $\boxed{\times} \boxed{10} \boxed{\wedge} \boxed{(-)} \boxed{34}$, using the $\boxed{(-)}$ key for the minus sign.

Desmos. Type 6.626*10^-34.

1.7 Periodic trends

Each trend comes from how many protons pull on the outer electrons and how far away those electrons are

General trend	Across →	Down ↓
Atomic radius	smaller	larger
Ionization energy	higher	lower
Electronegativity	higher	lower
Electron affinity energy released on gaining an electron (many exceptions)	more	less



Why the trends happen

Across a period, each element adds a proton while the outer electrons stay in the same shell, where they shield each other poorly. The pull grows, so atoms shrink and hold their electrons more tightly.

Down a group, each step adds a shell. The outer electrons sit farther out, behind more core electrons, so the pull weakens and atoms grow.

Atomic radii in picometers (pm), calculated values. Other tables define the radius differently and give other numbers, so exam questions print the radii they want you to compare. Don't compare these with the ionic radii below, which come from a different table.

□ s-block element □ p-block element

Ions with the same number of electrons



All six have 10 electrons, but more protons (7 to 13) pull them closer. Radii in pm, from crystals where each ion touches six others (four for N³⁻).

A cation is smaller than its atom.

Na⁺ has lost its only $n = 3$ electron, so its outer electrons are in $n = 2$.

An anion is larger than its atom.

Cl⁻ has one more electron but the same 17 protons, so its electrons repel more and spread out.

Comparing ion sizes

Li⁺ (76 pm) vs Na⁺ (102 pm): Na⁺'s outer electrons are in $n = 2$ and Li⁺'s in $n = 1$, so Na⁺'s are farther out and Na⁺ is larger.

K vs K⁺: K's outer electron is in $n = 4$, but the outer electrons of K⁺ are in $n = 3$, so K is larger.

Fe²⁺ vs Fe³⁺: both have 26 protons, but Fe²⁺ has one more electron, so more electron-electron repulsion, and it is larger.

Ions grow down a group too

Li⁺ 76 Na⁺ 102 K⁺ 138 Rb⁺ 152
F⁻ 133 Cl⁻ 181 Br⁻ 196 I⁻ 220

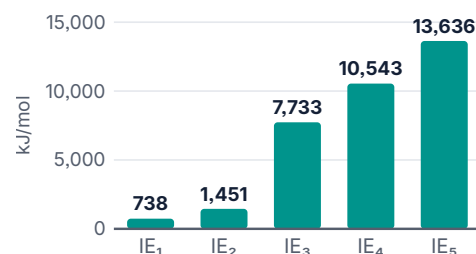
Ionic radii in pm. Each step down adds a shell.

Quick check

Rank by radius, largest first: Ca²⁺, Cl⁻, K⁺, S²⁻.

Answer: S²⁻ > Cl⁻ > K⁺ > Ca²⁺. All four have 18 electrons, and the protons rise from 16 to 20.

Successive ionization energies of Mg



IE₁ is the energy to remove the first electron, IE₂ the second. The jump at IE₃ shows two valence electrons: a third would come from the $n = 2$ core. So Mg forms Mg²⁺, never Mg³⁺.

Comparing how strongly two ions attract water

Compared with Na⁺, Mg²⁺ has the **greater charge**, so it attracts water more strongly. Mg²⁺ is also **smaller**, so the **distance between the ion and the O of a water molecule** is shorter, which makes the attraction stronger (Coulomb's law).

In Coulomb's law, r is the distance between the two attracting particles, not the radius of one ion.

Ionization energy, electronegativity, electron affinity

Ionization energy: the energy to remove one electron from each atom of a gas, $X(g) \rightarrow X^+(g) + e^-$.

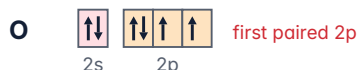
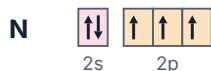
Electronegativity: how strongly an atom pulls on the electrons it shares in a bond. It follows the same trend as ionization energy, and fluorine has the highest.

Electron affinity: the energy released when a gaseous atom gains an electron. Cl releases more than F, because F is so small that its electrons repel an added one.

The two dips in ionization energy



Be's ionization energy is higher than B's (and Mg's than Al's). B's outer electron is in 2p, higher in energy and partly shielded by 2s, so it comes off more easily.



N's ionization energy is higher than O's (and P's than S's). O's fourth 2p electron shares an orbital, and its partner's repulsion makes it easier to remove.

Exam trap

- **Naming the trend instead of the reason.** "Ge is below Si" or "it's the periodic trend" earns nothing, and "Ge has more shells" isn't enough alone. Name both shells and link them to the pull: Ge's outer electrons are in $n = 4$ and Si's in $n = 3$, so Ge's are farther out and held less strongly.
- **Using "shielding" as the whole answer.** Say what differs: Cl has 17 protons and Na 11, with the same 10 core electrons, so Cl pulls harder.

1.8 Valence electrons and ionic compounds

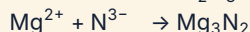
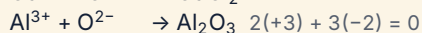
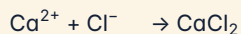
1	2	13	14	15	16	17	18
+1	+2	+3		-3	-2	-1	0
Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar
K	Ca	Ga	Ge	As	Se	Br	Kr
Rb	Sr	In	Sn	Sb	Te	I	Xe

Each group's usual ion charge is under its number; group 14 has none, and noble gases (0) form no ions. In periods 2 and 3 each ion has as many electrons as a noble gas: Na^+ , Mg^{2+} , O^{2-} and F^- all have neon's 10. Boron forms covalent compounds, not B^{3+} ions.

s block p block

Writing a formula from the charges

The charges must add up to zero, using the smallest whole numbers that do it:



Elements in the same group form similar formulas

NaCl , KCl , RbCl ; MgO , CaO , SrO ; H_2O , H_2S ; CCl_4 , SiCl_4 . Know one formula and you can predict the rest of its group.

Reactivity	Increases	Because
Metals (lose electrons)	down a group	lower ionization energy
Nonmetals (gain electrons)	up a group	stronger pull on electrons (F is the most reactive)

When a metal reacts with a nonmetal, electrons move from the metal to the nonmetal and ions form, giving an ionic compound. Two nonmetals share electrons instead (covalent bonds, Unit 2).

Quick check

Predict the formulas of the compounds of Sr with Br, and of Li with N.

Answer: Sr^{2+} and Br^- give SrBr_2 ; Li^+ and N^{3-} give Li_3N .

Finding the charge from ionization energies

Aluminum's first four ionization energies are 578, 1,817, 2,745 and 11,578 kJ/mol. The big jump after the third electron shows three valence electrons, so aluminum forms Al^{3+} .

What earns points on Unit 1 free-response questions

When the question asks you to...

Write this

Explain, "using principles of atomic structure," why X is larger or has a lower ionization energy than Y (different periods)

Write, for example: "X's outer electrons are in $n = 4$ and Y's in $n = 3$, so X's are farther from the nucleus and the Coulombic attraction is weaker." Name both atoms.

Make the same comparison within one period

The outer electrons are in the same shell, but the atom with more protons (and about the same shielding) pulls harder, so it is smaller and usually has the higher ionization energy (check for the two dips).

Compare an atom with its own ion, or two ions of one element

The nucleus is the same, so compare electrons: Na^+ has lost its whole outer shell, and Fe^{3+} has fewer electrons repelling each other than Fe^{2+} . Don't cite protons; they haven't changed.

Use Coulomb's law to compare how strongly two ions attract water

Compare the charges (larger charge, stronger attraction) and the distance between the two particles (smaller ion, closer approach). Write both comparisons.

Explain a dip in ionization energy

For B or Al, the electron removed is in a p subshell, higher in energy than s. For O or S, it shares an orbital and its partner repels it.

Identify an element from a photoelectron spectrum, or draw a missing peak

Label the peaks 1s, 2s, 2p and so on, turn heights into electrons and add them to get the atomic number. Draw a missing outermost peak at the lowest energy, as tall as its electron count.

Calculate moles, a mass, a percent or a formula

Show the setup with units, the formula's mole ratio as a factor, molar masses from the periodic table, and an answer with the right significant figures.

Say whether a lab error makes a result greater, less or the same

Trace the error to the mass you measured, then to the moles or percent you calculated, and state the direction.

Free-response questions about aluminum, worked through one point per row

The question	An answer that earns the point	How to earn it
Write the complete electron configuration of Al^{3+} .	$1s^2 2s^2 2p^6$	Show every subshell and superscript. Adding the atom's $3s^2 3p^1$ electrons loses the point.
Explain the jump from the third ionization energy (2,745 kJ/mol) to the fourth (11,578).	The fourth electron comes from $n = 2$, closer to the nucleus than the $n = 3$ valence electrons, so it is held much more strongly.	Name both shells and say the closer electron is held more strongly.
Explain why Al's first ionization energy is lower than Mg's.	Al's outer electron is in 3p, which is higher in energy than Mg's 3s and partly shielded by the 3s electrons, so it takes less energy to remove.	Name the 3p subshell. "Al has more protons" would predict a higher ionization energy, the wrong direction.
Explain why Al^{3+} is smaller than Al.	Al's outer electrons are in $n = 3$ and Al^{3+} 's in $n = 2$, closer to the same nucleus with 13 protons.	Compare the two shells; the protons don't change.
A 0.500 g sample of impure aluminum forms 0.850 g of Al_2O_3 . Find the percent Al.	$0.850 \div 101.96 = 0.008337$ mol Al_2O_3 ; $\times 2 = 0.01667$ mol Al; $\times 26.98 = 0.4498$ g; $\div 0.500$ g $\times 100 = 90.0\%$	Use 2 mol Al per mol Al_2O_3 , divide by the original sample's mass, and give three significant figures.
The oxide was still damp when weighed. Is the percent Al too high, too low or unchanged?	Too high: the mass of Al_2O_3 weighed was too high, so the moles of Al calculated were too high.	Give the direction and the chain from the mass to the answer.