ReactQuest — Flashcards

A Spark & Anvil printable flashcard deck. Quiz yourself on the ReactQuest cast's curriculum — one card at a time.

How to use these cards

Fold each sheet down the middle line so the answers are hidden. Read the **question** on the left, say your answer out loud, then **unfold** to check the right side. Testing yourself — then checking — is one of the most powerful ways to remember. Get one wrong? That's the best kind of practice: try it again tomorrow.

Want real flashcards? Cut along the fold line and the rows into cards — question on one side, answer on the other.

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1. Atomic structure — protons, neutrons, electrons, isotopes

✂ Question (fold →)	Answer
The charge of a proton is:	**positive (+1)** — Protons carry a +1 charge; electrons −1; neutrons 0.
The number of valence electrons largely determines an element's:	**chemical reactivity and bonding** — Chemistry is mostly the behavior of valence electrons — how atoms share or transfer them.
A sodium ion Na^+ has, compared to a neutral sodium atom:	**one fewer electron** — The + charge means it lost one electron; protons are unchanged.
Isotopes of an element differ in their number of:	**neutrons** — Isotopes share the proton count (same element) but differ in neutrons, so they have different masses.
Why is atomic structure the foundation for all of chemistry?	**the arrangement of electrons determines how atoms bond and react** — Every chemical behavior — bonding, reactivity, periodic trends — flows from how electrons are arranged around the nucleus.
Chlorine-35 and Chlorine-37 have the same:	**number of protons (17)** — Both are chlorine (17 protons); they differ in neutrons (18 vs 20) and mass.
The number of protons in a neutral sodium atom (atomic number 11) is:	**11** — The atomic number IS the proton count: 11.
An atom that GAINS an electron becomes a:	**negative ion (anion)** — Gaining an extra electron gives an excess of negative charge — an anion (−).
Average atomic mass on the periodic table is a weighted average of:	**the naturally occurring isotopes** — The listed atomic mass averages the isotope masses weighted by their natural abundance.
Which change would turn one element into a DIFFERENT element?	**changing the number of protons** — Only the proton count defines the element; changing electrons makes an ion, changing neutrons makes an isotope.
Electrons in the OUTERMOST shell are called:	**valence electrons** — Valence electrons are the outer-shell electrons that govern bonding and reactivity.
An atom that LOSES an electron becomes a:	**positive ion (cation)** — Losing a negative electron leaves an excess of positive protons — a cation (+).
Carbon-12 and Carbon-14 are isotopes because they both have 6 protons but:	**different numbers of neutrons (6 vs 8)** — Both are carbon (6 protons); 12 − 6 = 6 neutrons and 14 − 6 = 8 neutrons.
The MASS number of an atom equals the number of:	**protons plus neutrons** — Mass number counts the nucleons: protons + neutrons (electrons are too light to matter).
A neutral atom has equal numbers of:	**protons and electrons** — Neutral means the positive protons and negative electrons balance in number.
The atomic NUMBER of an element equals the number of:	**protons** — The atomic number is the proton count — it defines which element an atom is.
The MASS of an atom is	**nucleus** — Protons and neutrons (in the nucleus) carry nearly all the mass;

concentrated in its:	electrons are ~1/1836 as massive.
An atom with 8 protons, 8 neutrons, and 8 electrons is:	**a neutral oxygen atom** — 8 protons = oxygen; equal protons and electrons = neutral.
A common misconception is that an isotope is a different ELEMENT. In fact isotopes are:	**the same element with a different mass (neutron count)** — Same protons = same element; only the neutron count (and mass) differs. Not all isotopes are radioactive.
The nucleus of an atom contains the:	**protons and neutrons** — The dense central nucleus holds the protons and neutrons; electrons orbit outside it.
The number of electrons that can occupy the first energy level is:	**2** — The first (innermost) energy level holds a maximum of 2 electrons.
An element's identity is found on the periodic table using its:	**atomic number (proton count)** — The periodic table is ordered by atomic number, which uniquely identifies each element.
An ION forms when an atom:	**gains or loses electrons, becoming charged** — Ions form by electron transfer; the atom keeps its protons (still the same element) but gains a charge.
The three subatomic particles are:	**protons, neutrons, and electrons** — An atom is built from protons (+), neutrons (neutral) in the nucleus, and electrons (−) around it.
Electrons occupy regions around the nucleus called:	**energy levels (shells / orbitals)** — Electrons occupy energy levels/orbitals surrounding the nucleus, not the nucleus itself.

2. Periodic trends — organization and predictable patterns

✂ Question (fold →)	Answer
The periodic table is arranged in order of increasing:	**atomic number** — Elements are ordered by atomic number (proton count), which produces the repeating patterns.
The most reactive nonmetals are found:	**in the upper right (Group 17, halogens)** — Halogens need just one electron to fill their shell, making them very reactive nonmetals.
Group 17 elements (halogens) tend to form ions with a charge of:	**−1** — Gaining one electron to complete the shell gives halogens a −1 charge.
Horizontal rows in the periodic table are called:	**periods** — Rows are periods; across a period, properties change gradually.
Vertical columns in the periodic table are called:	**groups (or families)** — Columns are groups; elements in a group share similar chemical behavior (same valence count).
The most reactive metals are found:	**in the lower left (Group 1, alkali metals)** — Alkali metals lose their single valence electron easily, making them highly reactive.
Ionization energy (energy to remove an electron) generally INCREASES:	**left to right across a period** — A stronger nuclear pull to the right holds electrons more tightly, raising ionization energy.
A key trend misconception is that atoms get BIGGER left-to-right across a period. Actually they get:	**smaller (increasing nuclear charge pulls electrons in)** — Despite adding electrons, the growing proton count pulls the shell tighter — atoms shrink across a period.
Moving DOWN Group 1, the metals become:	**more reactive** — Down Group 1 the outer electron is farther from the nucleus and lost more easily, raising reactivity.
Why did organizing elements by periodic trends transform chemistry?	**it revealed predictable patterns and even predicted undiscovered elements** — Mendeleev's arrangement exposed repeating patterns and successfully predicted elements not yet found — a triumph of scientific reasoning.
As electronegativity decreases, an element becomes more:	**metallic (electron-losing)** — Low electronegativity means weak attraction for electrons — metallic, electron-donating behavior.
The periodic table's repeating patterns arise because elements in a column share:	**the same valence electron arrangement** — Recurring valence arrangements down each cycle produce the periodic repetition of properties.
The noble gases (Group 18) are very UNreactive because they have:	**a full valence shell** — A complete outer shell makes noble gases stable and reluctant to bond.
Electronegativity (attraction for shared electrons) is HIGHEST:	**in the upper right (excluding noble gases), e.g. fluorine** — Electronegativity rises up and to the right; fluorine is the most electronegative element.
Elements in Group 18 (noble gases) are grouped together because they all have:	**a full (stable) valence shell** — Noble gases share a complete valence shell, making them similarly unreactive.

Group 1 elements (alkali metals) tend to form ions with a charge of:	**+1** — Losing their single valence electron gives alkali metals a +1 charge.
Atomic radius generally INCREASES as you move:	**down a group** — Down a group each element adds an electron shell, making the atom larger.
Which element is more reactive as a metal: sodium (Na) or magnesium (Mg)?	**sodium (it loses its single valence electron more easily)** — Sodium (Group 1) sheds one electron more readily than magnesium (Group 2), so it's more reactive.
Metals are generally located on the __ of the periodic table:	**left and center** — Metals occupy the left and center; nonmetals sit in the upper right, with metalloids on the stair-step.
The transition metals occupy the:	**center block of the table (Groups 3–12)** — Transition metals fill the central d-block of the periodic table.
Metalloids (like silicon) are found:	**along the stair-step line between metals and nonmetals** — Metalloids straddle the metal/nonmetal boundary and show intermediate properties.
Predicting reactivity from position illustrates the power of the periodic table because it lets you:	**predict behavior without testing every element** — The table encodes trends so you can predict radius, reactivity, and bonding straight from an element's location.
Atomic radius generally DECREASES as you move:	**left to right across a period** — Across a period the growing nuclear charge pulls the same shell in tighter, shrinking the atom.
Potassium (K) is below sodium (Na) in Group 1, so potassium's atomic radius is:	**larger than sodium's** — Down a group atoms gain shells, so potassium (period 4) is larger than sodium (period 3).
Elements in the same GROUP have the same number of:	**valence electrons** — Shared valence-electron count is why a group's elements react similarly.

3. Chemical bonding — ionic, covalent, and Lewis structures

✂ Question (fold →)	Answer
A metallic bond involves:	**a 'sea' of shared electrons among metal cations** — In metals, valence electrons delocalize into a shared sea around positive metal ions — explaining conductivity and malleability.
Ionic compounds tend to have __ melting points:	**high** — Strong lattice attractions give ionic compounds high melting points.
To predict whether a bond is ionic or covalent, compare the atoms':	**electronegativity difference** — A large electronegativity difference → ionic; a small one → covalent (polar or nonpolar).
A COVALENT bond forms by:	**sharing electrons between atoms** — Covalent bonds share pairs of electrons between atoms.
A common misconception is that COVALENT bonds transfer electrons. Actually covalent bonds:	**share electrons** — Sharing (not transfer) is the defining feature of covalent bonding; transfer is ionic.
Water (H_2O) is a polar molecule because:	**oxygen pulls the shared electrons more strongly than hydrogen** — Oxygen's higher electronegativity makes the O–H bonds polar, and the bent shape makes the whole molecule polar.
The octet rule says atoms tend to bond so as to have:	**eight valence electrons (a full outer shell)** — Most main-group atoms bond to reach a stable octet of eight valence electrons.
Which bond type conducts electricity when DISSOLVED in water?	**ionic** — Dissolved ionic compounds release mobile ions that carry current; molecular covalent compounds usually don't.
Predicting a compound's properties from its bond type illustrates that:	**bonding determines macroscopic properties like melting point and conductivity** — Whether a substance is a hard high-melting crystal or a soft molecular gas traces directly to its bond type.
A single covalent bond consists of:	**one shared pair of electrons** — A single covalent bond shares exactly one pair (two electrons) between the atoms.
The number of covalent bonds a carbon atom typically forms is:	**4** — Carbon has 4 valence electrons and forms 4 covalent bonds to complete its octet — the basis of organic chemistry.
A chemical bond forms because atoms:	**achieve a more stable electron arrangement (often a full shell)** — Atoms bond to reach a lower-energy, more stable configuration — usually a filled valence shell.
A Lewis (dot) structure shows an atom's:	**valence electrons as dots around the symbol** — Lewis structures depict valence electrons (dots) to track bonding.
A common bonding misconception is that ionic compounds exist as 'molecules.' Actually they form:	**a repeating lattice of ions (a crystal), not discrete molecules** — Ionic compounds are 3-D lattices of alternating ions; 'a molecule of NaCl' is a misconception.
Why is understanding bonding central to chemistry?	**it explains how and why atoms combine to form all substances** — Every compound, reaction, and material property depends on how atoms bond — the core of chemical behavior.
A nonpolar covalent bond forms	**equally (identical or very similar electronegativity)** — Equal sharing (e.g.

when atoms share electrons:	in O_2) gives a nonpolar bond with no partial charges.
In NaCl, sodium becomes Na^+ and chlorine becomes Cl^- because:	**sodium transferred its one valence electron to chlorine** — Na loses 1 electron (→ Na^+) and Cl gains it (→ Cl^-); opposite charges attract ionically.
Diamond and graphite are both made of carbon but differ in properties because of their:	**different bonding arrangements (structure)** — Same element, different bonding network — so diamond is hard and graphite is soft/slippery.
A triple bond (as in N_2) shares:	**three pairs of electrons (six electrons)** — A triple bond shares three electron pairs, making a very strong bond (as in nitrogen gas).
A polar covalent bond forms when the two atoms:	**share electrons UNequally (different electronegativities)** — Unequal electronegativity pulls the shared electrons toward one atom, creating partial charges (polar bond).
A double bond consists of:	**two shared pairs of electrons** — A double bond shares two electron pairs (four electrons) between two atoms.
An IONIC bond forms by:	**transferring electrons from one atom to another** — Ionic bonds transfer electrons, creating oppositely charged ions that attract.
Covalent bonds typically form between:	**two nonmetals** — Two nonmetals share electrons, forming a covalent (molecular) bond.
Ionic bonds typically form between:	**a metal and a nonmetal** — A metal (loses electrons) plus a nonmetal (gains them) makes an ionic compound.
The chemical formula H_2O tells you a molecule contains:	**2 hydrogen atoms and 1 oxygen atom** — Subscripts give the atom counts: 2 H and 1 O (subscript 1 is implied).

4. Naming compounds and writing chemical formulas

✂ Question (fold →)	Answer
A polyatomic ion like sulfate (SO_4^{2-}) is:	**a charged group of atoms that acts as a single unit** — Polyatomic ions are covalently bonded atom groups carrying an overall charge, treated as one unit in formulas.
An -ATE ending on a polyatomic ion (like sulfate, nitrate) usually indicates the ion:	**contains oxygen (an oxyanion)** — Oxyanions (containing oxygen) commonly end in -ate (or -ite for one fewer oxygen).
Correct chemical names and formulas matter because they:	**communicate an exact composition unambiguously worldwide** — A precise name/formula tells any chemist exactly what substance is meant — the shared language of chemistry.
K_2O is named:	**potassium oxide** — Potassium is a fixed +1 metal (no Roman numeral, no prefixes for ionic): potassium oxide.
The difference between CO and CO_2 is:	**the number of oxygen atoms (1 vs 2), giving different compounds** — CO (carbon monoxide) and CO_2 (carbon dioxide) are entirely different substances — subscripts matter.
Writing a formula correctly is a prerequisite for stoichiometry because:	**the subscripts and coefficients drive every mole calculation** — A wrong formula cascades into wrong mole ratios and wrong quantities — accurate formulas are load-bearing.
The formula $CaCl_2$ contains how many chlorine atoms?	**2** — The subscript 2 on Cl means two chlorine atoms per formula unit.
The prefix 'mono-' means:	**one** — mono- = 1, di- = 2, tri- = 3, tetra- = 4.
Fe_2O_3 is named:	**iron(III) oxide** — Each O is −2 (total −6); two Fe share +6, so each Fe is +3 → iron(III) oxide.
The compound of Na^+ and O^{2-} has the formula:	**Na_2O** — Two Na^+ (+1 each) balance one O^{2-} (−2): Na_2O.
For COVALENT (molecular) compounds, Greek prefixes indicate:	**the number of each atom (mono-, di-, tri-…)** — Molecular compounds use prefixes to count atoms: CO_2 = carbon dioxide (di- = 2 oxygens).
Why is nomenclature the 'grammar' of chemistry?	**it lets you translate reliably between names, formulas, and real substances** — Just as grammar structures language, nomenclature rules let you move confidently between a name, its formula, and the actual compound.
CO_2 is named:	**carbon dioxide** — One carbon (mono- often dropped on the first element), two oxygens (di-): carbon dioxide.
For a transition metal with more than one possible charge, the charge is shown with:	**a Roman numeral, e.g. iron(III)** — Variable-charge metals use a Roman numeral to specify the ion charge: iron(II) vs iron(III).
To write a formula from a name like 'aluminum oxide,' you:	**find each ion's charge (Al^{3+}, O^{2-}) and balance them (Al_2O_3)** — Al^{3+} and O^{2-} balance as Al_2O_3 (cross the charges to the smallest whole-number ratio).

The formula for an ionic compound is written so that the total charge is:	**zero (charges balance)** — Ionic formulas use the smallest whole-number ratio of ions that makes the compound electrically neutral.
The compound of Ca^{2+} and NO_3^- (nitrate) is written:	**$Ca(NO_3)_2$** — One Ca^{2+} needs two nitrate ions; parentheses keep the nitrate intact: $Ca(NO_3)_2$.
Naming H_2O by rules would give 'dihydrogen monoxide,' but its common name is:	**water** — Some compounds keep common names; H_2O is universally 'water.'
The name of NaCl is:	**sodium chloride** — Metal first (sodium), nonmetal with -ide (chloride): sodium chloride.
The nonmetal (anion) in an ionic compound gets the ending:	**-ide** — A simple anion ends in -ide: chlorine → chloride, oxygen → oxide.
When naming an ionic compound, the metal (cation) is named:	**first, with its element name unchanged** — The metal keeps its name and comes first: sodium chloride (Na then Cl).
In a chemical formula, a SUBSCRIPT tells you the number of:	**atoms of the element it follows** — A subscript counts the atoms of the immediately preceding element (H_2O has 2 H).
The ionic compound of Mg^{2+} and Cl^- has the formula:	**$MgCl_2$** — One Mg^{2+} needs two Cl^- to balance the charge: $MgCl_2$.
The ionic compound of Na^+ and Cl^- has the formula:	**NaCl** — +1 and −1 balance one-to-one, so the formula is NaCl.
A common formula-writing misconception is to change SUBSCRIPTS to balance charge. Correct practice is to:	**adjust the RATIO of ions (crossing charges), keeping each ion's identity** — You balance charge by choosing how many of each ion, not by altering a polyatomic ion's internal subscripts.

5. The mole and stoichiometry — quantitative chemistry

✂ Question (fold →)	Answer
You balance a chemical equation by adjusting:	**coefficients (the big numbers in front)** — Only coefficients may change; altering subscripts would change the substances themselves.
Why is stoichiometry the quantitative heart of chemistry?	**it lets you predict and control exact amounts in any reaction** — From medicine doses to industrial yields, stoichiometry turns balanced equations into precise, controllable quantities.
The coefficients in a balanced equation give the ratio of:	**moles of each substance** — Coefficients are mole ratios — the bridge for stoichiometric conversions.
To find moles of product from moles of a reactant, you multiply by the:	**mole ratio from the balanced equation** — Multiply the known moles by the ratio of product-to-reactant coefficients.
A reaction that produces a GAS may APPEAR to lose mass in an open container because:	**the gas escapes and isn't weighed — mass is still conserved in a closed system** — The escaping gas carries mass away; sealed, the total mass would be unchanged. This is a classic misconception.
A common balancing misconception is to change SUBSCRIPTS. This is wrong because it:	**changes the identity of the compound** — Changing H_2O to H_2O_2 makes a different substance; you must balance with coefficients instead.
To convert grams of a substance to moles, you:	**divide by the molar mass** — moles = grams ÷ molar mass (g/mol).
The mole concept is powerful because it:	**connects the invisible particle count to a measurable mass in the lab** — The mole bridges Johnstone's submicroscopic (particles) and macroscopic (grams) levels — letting us count atoms by weighing.
A common mole misconception is that a mole is a fixed MASS. In fact a mole is:	**a fixed NUMBER of particles (mass differs per substance)** — A mole of any substance has the same particle count but a different mass (its molar mass).
For $2H_2 + O_2 \rightarrow 2H_2O$, how many moles of water form from 4 mol H_2 (with excess O_2)?	**4 moles** — The H_2:H_2O ratio is 1:1, so 4 mol H_2 → 4 mol H_2O.
A balanced chemical equation must obey:	**conservation of mass (same atoms on both sides)** — Balancing ensures every atom is accounted for on both sides — matter is neither created nor destroyed.
Avogadro's number is defined as the number of particles in:	**one mole of any substance** — 6.022×10^{23} is the count of particles in exactly one mole.
Balancing the equation before doing stoichiometry is essential because:	**the coefficients ARE the mole ratios the whole calculation depends on** — An unbalanced equation gives wrong mole ratios, so every downstream amount is wrong.
The general stoichiometry road-map is: grams A → moles A → moles B → grams B, using:	**molar masses and the mole ratio** — Convert mass to moles, cross the mole bridge (balanced ratio), then convert back to mass.

Another mole misconception: 'equal masses of two substances contain equal moles.' This is:	**false — equal masses give different moles (different molar masses)** — 12 g of carbon is 1 mol, but 12 g of water is 0.67 mol — equal mass ≠ equal moles.
Stoichiometry lets a chemist predict:	**how much product a given amount of reactant will make** — Stoichiometry converts between amounts of reactants and products using the balanced equation.
The molar mass of CO_2 (C = 12, O = 16) is:	**44 g/mol** — 12 + 16 + 16 = 44 g/mol.
Conservation of mass predicts that if 10 g of reactants fully react, the products weigh:	**10 g (mass is conserved)** — Atoms are rearranged, not created or destroyed, so total mass is unchanged: 10 g in, 10 g out.
In the balanced equation $2H_2 + O_2 \rightarrow 2H_2O$, the mole ratio of H_2 to H_2O is:	**2 : 2 (i.e. 1 : 1)** — Coefficients give the ratio directly: 2 H_2 to 2 H_2O = 1:1.
A mole is a unit that counts a specific NUMBER of particles:	**6.022×10^{23} (Avogadro's number)** — A mole is a COUNT — Avogadro's number of particles — not a mass.
To convert moles to number of particles, you:	**multiply by Avogadro's number (6.022×10^{23})** — particles = moles × 6.022 × 10^{23}.
How many moles are in 36 g of water (molar mass 18 g/mol)?	**2 moles** — 36 g ÷ 18 g/mol = 2 mol.
The number of particles in 2 moles of a substance is:	**$2 \times 6.022 \times 10^{23}$** — Multiply moles by Avogadro's number: 2 × 6.022 × 10^{23} particles.
The coefficient in front of a formula in a balanced equation multiplies:	**every atom in that formula** — A coefficient applies to the whole formula unit — every atom it contains.
The molar mass of an element (grams per mole) is numerically equal to its:	**average atomic mass on the periodic table** — The atomic mass in amu equals the molar mass in grams per mole (e.g. carbon ≈ 12 g/mol).

6. Limiting reagent and percent yield

✂ Question (fold →)	Answer
If a reaction has a low percent yield, possible causes include:	**side reactions, incomplete reaction, or product loss during recovery** — Real yields drop due to competing reactions, equilibrium limits, and handling losses.
To find the theoretical yield, you do a stoichiometry calculation starting from:	**the limiting reagent** — The limiting reagent sets the maximum product, so the theoretical-yield calculation begins there.
Percent yield is calculated as:	**(actual yield ÷ theoretical yield) × 100%** — Percent yield = actual/theoretical × 100 — how close you got to the ideal.
Once the limiting reagent is consumed, the reaction:	**stops making product (excess reagent remains)** — Without the limiting reagent, no more product can form even if excess reactant is present.
The ACTUAL yield is:	**the amount of product actually obtained in the experiment** — Actual yield is what you really collect — usually less than theoretical due to losses.
The THEORETICAL yield is:	**the maximum product predicted by stoichiometry from the limiting reagent** — Theoretical yield is the ideal maximum calculated from the limiting reagent.
The EXCESS reagent is the reactant that:	**remains left over after the reaction stops** — The excess reagent is not fully consumed; some remains once the limiting reagent is gone.
Given 3 mol H_2 and 3 mol O_2 for $2H_2 + O_2 \rightarrow 2H_2O$, the limiting reagent is:	**H_2** — The ratio needs 2 H_2 per 1 O_2; 3 mol H_2 pairs with only 1.5 mol O_2, so H_2 runs out first.
The amount of product is determined by:	**the limiting reagent** — Once the limiting reagent is gone, no more product forms — it sets the maximum.
A key habit before a limiting-reagent problem is to:	**write and balance the equation, then convert everything to moles** — Balancing and converting to moles are the prerequisites that prevent the mass-comparison error.
In that reaction, how much O_2 is left over (started with 3 mol)?	**1.5 mol O_2 (excess)** — 3 mol H_2 uses 1.5 mol O_2, leaving 3 − 1.5 = 1.5 mol O_2 in excess.
The LIMITING reagent in a reaction is the reactant that:	**runs out first, stopping the reaction** — The limiting reagent is used up first and thereby caps how much product can form.
Percent yield matters in industry because it measures:	**how efficiently reactants are converted to product** — Percent yield tracks reaction efficiency — a key economic and process metric.
Why is the limiting-reagent concept essential to real chemistry?	**reactions in the real world rarely use exact amounts, so one reactant almost always limits the product** — Because real mixtures are rarely perfectly proportioned, identifying the limiting reagent is what predicts actual output.
A percent yield GREATER than 100% usually indicates:	**experimental error (e.g. impure or wet product)** — Since you can't make more than theoretical, over-100% points to contamination or leftover solvent — not created matter.
A common limiting-reagent	

misconception is that the reactant with the smaller MASS is limiting. Actually you must compare:	**moles (using the balanced mole ratio), not masses** — Limiting reagent is decided by moles relative to the balanced ratio — a small mass can still be in excess.
Why do industrial chemists often deliberately use an EXCESS of one reactant?	**to ensure the more valuable reactant is fully consumed** — Excess of a cheap reactant drives the expensive/limiting reagent to react completely, maximizing use of the valuable material.
To find the limiting reagent, you first convert each reactant to:	**moles** — Convert given masses to moles, then compare against the balanced mole ratio.
Comparing available moles to the balanced ratio is the key step because it reveals:	**which reactant is in short supply relative to what's needed** — The mole-to-ratio comparison exposes the reactant that will run out first.
The percent yield can never exceed 100% because:	**you cannot make more product than stoichiometry allows** — Theoretical yield is the maximum; a value over 100% signals error, not extra product.
If the theoretical yield is 50 g and you obtain 40 g, the percent yield is:	**80%** — (40 ÷ 50) × 100 = 80%.
If two reactants are present in EXACTLY the stoichiometric ratio:	**both are fully consumed (neither is in excess)** — Perfect stoichiometric proportions mean both reactants run out together — no excess.
Identifying the limiting reagent requires the balanced equation because it provides:	**the mole ratio to compare the reactants against** — You compare available moles to the balanced mole ratio to see which reactant runs short.
Predicting the limiting reagent before running a reaction lets a chemist:	**plan reactant amounts and predict the maximum product** — Knowing the limiting reagent in advance guides how much of each reactant to use and what yield to expect.
The reactant left over at the end of a reaction is the:	**excess reagent** — Whatever remains after the reaction stops is the excess reagent.

7. Gas laws — pressure, volume, temperature relationships

✂ Question (fold →)	Answer
Charles's Law states that, at constant pressure, volume and temperature are:	**directly proportional (T up, V up)** — At fixed pressure, heating a gas expands it: $V \propto T$ (in kelvin).
Gay-Lussac's Law relates, at constant volume:	**pressure and temperature (directly proportional)** — At fixed volume, heating a gas raises its pressure: $P \propto T$.
In PV = nRT, if n (moles of gas) increases at constant T and V, the pressure:	**increases** — More gas particles in the same volume means more collisions — higher pressure.
The Combined Gas Law is:	$\mathbf{P_1V_1/T_1 = P_2V_2/T_2}$ — The combined law links P, V, and T between two states of a fixed amount of gas.
A common gas-law misconception is using CELSIUS in calculations. This fails because:	**the proportions require an absolute (kelvin) scale with a true zero** — Direct proportionality to temperature only works from absolute zero, so kelvin is mandatory.
To convert Celsius to kelvin, you:	**add 273 (K = °C + 273)** — K = °C + 273 (more precisely 273.15).
The Ideal Gas Law is:	**PV = nRT** — PV = nRT links pressure, volume, moles (n), and temperature via the gas constant R.
Two identical balloons, one warm and one cold, at the same pressure: the warmer balloon is:	**larger (Charles's law)** — At constant pressure, warmer gas expands to a larger volume ($V \propto T$).
Why are the gas laws a cornerstone of chemistry and engineering?	**they let us predict and control gas behavior in reactions, engines, and the atmosphere** — From respiration to combustion engines to weather, the gas laws model and predict how gases behave under changing conditions.
Real gases deviate most from ideal behavior at:	**high pressure and low temperature** — Under high pressure/low temperature, particle volume and attractions matter, so real gases deviate from PV = nRT.
In PV = nRT, doubling the number of moles n while holding P and T constant:	**doubles the volume** — At constant P and T, volume is proportional to moles, so doubling n doubles V.
If the volume of a gas is doubled at constant temperature, its pressure:	**halves** — $P \propto 1/V$, so doubling volume halves the pressure.
Avogadro's Law states that equal VOLUMES of gases at the same T and P contain:	**equal numbers of molecules** — At the same temperature and pressure, equal gas volumes hold equal molecule counts, regardless of the gas.
If you heat a sealed, rigid container of gas, the pressure will:	**increase (Gay-Lussac's law)** — At constant volume, higher temperature means more forceful, frequent collisions — higher pressure.

In all gas-law calculations, temperature MUST be in:	**kelvin** — Gas laws require absolute temperature (kelvin); Celsius can be zero or negative and breaks the proportions.
If the kelvin temperature of a gas doubles at constant pressure, its volume:	**doubles** — $V \propto T$: doubling kelvin temperature doubles the volume.
Kinetic Molecular Theory says gas pressure comes from:	**gas particles colliding with the container walls** — Pressure is the cumulative force of countless particle–wall collisions.
Absolute zero (0 K) is the temperature at which particle motion is:	**at a theoretical minimum** — At 0 K, particle motion reaches its theoretical minimum — the basis of the kelvin scale.
In PV = nRT, the symbol 'n' represents the:	**number of moles of gas** — n is the amount of gas in moles.
The gas constant R connects the units in PV = nRT so that the equation:	**balances dimensionally (e.g. 0.0821 L·atm/mol·K)** — R is a fixed proportionality constant whose value depends only on the chosen units.
Using the gas laws, you can PREDICT a gas's response to a change because:	**P, V, T, and n are linked by a fixed mathematical relationship** — The gas laws quantify how altering one variable forces the others to respond — enabling precise prediction.
If a gas is compressed to half its volume at constant temperature, its pressure:	**doubles** — $P \propto 1/V$: halving V doubles P.
Boyle's Law states that, at constant temperature, pressure and volume are:	**inversely proportional (P up, V down)** — At fixed temperature, squeezing a gas into a smaller volume raises its pressure: $P \propto 1/V$.
At STP (standard temperature and pressure), one mole of any ideal gas occupies:	**22.4 liters** — One mole of an ideal gas fills 22.4 L at STP (0 °C, 1 atm).
Raising the temperature of a gas increases the particles':	**average kinetic energy (they move faster)** — Temperature is a measure of average kinetic energy; hotter gas particles move faster.

8. Solutions and concentration — molarity and dilution

✂ Question (fold →)	Answer
In salt water, the SOLUTE is:	**the salt** — The solute is the dissolved substance (salt); the solvent is what dissolves it (water).
The dilution equation is:	**$M_1V_1 = M_2V_2$** — Because moles of solute stay constant, $M_1V_1 = M_2V_2$ relates the concentrations and volumes.
Diluting 100 mL of 2 M solution to 200 mL gives a new concentration of:	**1 M** — $M_2 = M_1V_1/V_2 = (2)(100)/200 = 1$ M.
The concentration unit 'molarity' depends on the volume of:	**the whole solution (not just the solvent)** — Molarity uses liters of final solution, which differs slightly from the solvent volume added.
Solubility of most SOLIDS in water generally __ as temperature increases:	**increases** — Warmer water usually dissolves more solid solute (e.g. more sugar in hot tea).
How many moles of solute are in 0.5 L of a 3 M solution?	**1.5 mol** — moles = M × V = 3 × 0.5 = 1.5 mol.
A solution is a homogeneous mixture of a:	**solute dissolved in a solvent** — A solution uniformly disperses a solute (dissolved substance) throughout a solvent (dissolving medium).
A SATURATED solution is one that:	**holds the maximum solute that will dissolve at that temperature** — A saturated solution contains the maximum dissolvable solute; extra solute won't dissolve.
Preparing a solution of a precise molarity is important in the lab because:	**reaction amounts depend on the moles of solute delivered** — Accurate molarity ensures you deliver the intended moles of reactant — essential for reproducible chemistry.
When you dilute a solution, the NUMBER OF MOLES of solute:	**stays the same** — Dilution adds only solvent, so the amount of solute (moles) is unchanged — only concentration drops.
MOLARITY (M) is defined as:	**moles of solute per liter of solution** — Molarity = moles of solute ÷ liters of solution (mol/L).
A solution that can still dissolve more solute is called:	**unsaturated** — An unsaturated solution holds less than the maximum, so more solute can still dissolve.
Why must temperature be specified when stating solubility?	**solubility changes with temperature** — Because how much solute dissolves depends on temperature, a solubility value is meaningless without it.
A common concentration misconception is confusing 'concentrated' with 'strong.' Concentrated means:	**a lot of solute is dissolved (high molarity)** — Concentration is about how much solute is dissolved; 'strong' (for acids) is about how fully it ionizes — different ideas.
To find moles of solute from molarity and volume, you:	**multiply molarity by volume in liters (moles = M × V)** — moles = molarity (mol/L) × volume (L).
An electrolyte is a solute that, when dissolved:	**produces ions and conducts electricity** — Electrolytes (like ionic salts) dissociate into mobile ions that carry current in solution.

The 'like dissolves like' rule means:	**polar solvents dissolve polar/ionic solutes; nonpolar dissolves nonpolar** — Solubility is favored when solute and solvent have similar polarity (water dissolves salt; oil dissolves grease).
If you EVAPORATE some solvent from a solution, the concentration:	**increases** — Removing solvent leaves the same solute in less volume — higher concentration.
To predict a dilution's outcome, you use $M_1V_1 = M_2V_2$ because it captures that:	**the moles of solute are conserved while volume changes** — The equation encodes conservation of solute (moles) — the reason concentration and volume trade off predictably.
Why are solutions central to practical chemistry?	**most reactions, biology, and industrial processes happen in solution** — Blood, oceans, medicines, and most reactions occur in solution — controlling concentration is fundamental to real chemistry.
A solution with 2 moles of solute in 1 liter of solution has a molarity of:	**2 M** — M = moles/liters = 2/1 = 2 M.
If you add solute to a solution without adding solvent, the concentration:	**increases** — More solute in the same volume raises the concentration.
To make a solution MORE dilute, you:	**add more solvent** — Adding solvent spreads the same solute through more volume, lowering the concentration.
Molarity has units of:	**moles per liter (mol/L)** — Molarity is measured in mol/L (also written M).
Dissolving is faster when the solvent is stirred because stirring:	**brings fresh solvent into contact with the solute** — Stirring moves saturated solvent away and fresh solvent in, speeding dissolving.

9. Reaction types — recognizing patterns of chemical change

✂ Question (fold →)	Answer
$2H_2 + O_2 \rightarrow 2H_2O$ is an example of a:	**synthesis (combination) reaction** — Two reactants (H_2 and O_2) combine into a single product (water) — synthesis.
Complete combustion of a hydrocarbon always produces:	**carbon dioxide and water** — Burning a hydrocarbon fully in oxygen yields CO_2 and H_2O.
A SINGLE-replacement reaction has the form:	**A + BC → AC + B (one element replaces another)** — A more reactive element displaces another from its compound.
$2H_2O \rightarrow 2H_2 + O_2$ is an example of a:	**decomposition reaction** — One compound (water) breaks into two elements — decomposition.
$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ is classified as:	**combustion** — A hydrocarbon burning in oxygen to give CO_2 and water is combustion.
Why is recognizing reaction types foundational for a chemist?	**it turns an unfamiliar reaction into a predictable pattern with known products** — Pattern recognition lets chemists anticipate products and energy behavior across countless reactions from a few templates.
Classifying a reaction as endothermic or exothermic is SEPARATE from classifying it as synthesis/decomposition because:	**the type describes the pattern; endo/exo describes the energy flow** — Reaction TYPE (structural pattern) and energy direction (endo/exo) are independent descriptors of the same reaction.
Recognizing a reaction TYPE helps a chemist:	**predict the products of a reaction** — Knowing the pattern lets you predict what products will form.
A COMBUSTION reaction involves a fuel reacting with:	**oxygen, releasing energy (often producing CO_2 and H_2O)** — Combustion burns a fuel in oxygen, typically yielding carbon dioxide, water, and energy.
A DECOMPOSITION reaction has the form:	**AB → A + B (one reactant breaks into two)** — Decomposition breaks one compound into simpler substances.
Predicting products from a reaction type, then balancing, is a core chemistry skill because it lets you:	**write a complete, correct equation for an unfamiliar reaction** — Recognizing the type predicts the products, and balancing completes a correct equation ready for stoichiometry.
The reaction of an acid with a base to form water and a salt is a type of:	**double-replacement (neutralization) reaction** — Acid + base swap partners to make water and a salt — a double-replacement neutralization.
A synthesis and a decomposition reaction are related in that they are:	**essentially reverses of each other** — Synthesis combines into one product; decomposition breaks one apart — opposite directions of the same idea.
A SYNTHESIS (combination) reaction has the form:	**A + B → AB (two reactants form one product)** — In synthesis, two or more substances combine into a single product.
The reaction $NaOH + HCl \rightarrow NaCl + H_2O$ is best classified as:	**double replacement (neutralization)** — An acid and base exchange partners to form water and a salt — double replacement.
$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$ is an example	**single-replacement reaction** — Zinc replaces copper in the sulfate

of a:	compound — single replacement.
A reaction where a compound gains oxygen or a metal rusts is broadly a type of:	**oxidation reaction (a redox process)** — Gaining oxygen (rusting, combustion) is oxidation — part of the redox family.
A single reaction can belong to MORE than one category, for example combustion is also a type of:	**oxidation (redox) reaction** — Combustion is both a combustion reaction and a redox (oxidation) reaction — categories can overlap.
In a double-replacement reaction, a solid product that forms and settles out is called a:	**precipitate** — An insoluble solid formed in a double-replacement reaction is a precipitate.
Predicting whether a single-replacement reaction will occur requires checking:	**whether the free element is more reactive than the one it would replace** — Use the activity series: a more reactive free element displaces a less reactive one; otherwise no reaction.
A more reactive metal can displace a less reactive metal from solution — this is predicted by the:	**activity series** — The activity series ranks metals by reactivity, predicting which single-replacement reactions occur.
In a single-replacement reaction, whether it proceeds depends on the:	**relative reactivity of the elements (activity series)** — A more reactive free element will displace a less reactive one; the activity series predicts this.
Balancing a reaction is required regardless of its type because:	**conservation of mass applies to every reaction** — Every reaction must conserve atoms, so balancing applies to all reaction types.
The general clue that a double-replacement reaction 'went' is the formation of:	**a precipitate, a gas, or water** — Double-replacement reactions proceed when they form an insoluble precipitate, a gas, or water.
A DOUBLE-replacement reaction has the form:	**AB + CD → AD + CB (partners switch)** — Two compounds exchange partners, often forming a precipitate, gas, or water.

10. Kinetics — reaction rate and collision theory

✂ Question (fold →)	Answer
Enzymes in living cells act as:	**biological catalysts** — Enzymes are biological catalysts that lower activation energy for life's reactions.
Increasing TEMPERATURE speeds up a reaction because particles:	**move faster and collide more often with more energy** — Higher temperature raises collision frequency and the fraction of collisions exceeding the activation energy.
A reaction generally goes faster when reactant particles collide:	**more frequently and with more energy** — Increasing collision frequency and energy raises the rate.
Collision theory unifies the rate factors because each factor changes:	**the frequency or effectiveness of collisions** — Temperature, concentration, surface area, and catalysts all work by altering how often and how effectively particles collide.
The minimum energy needed for a reaction to occur is the:	**activation energy** — Activation energy is the energy barrier that colliding particles must overcome to react.
A common kinetics misconception is that a catalyst is 'used up.' In fact a catalyst:	**is regenerated and can be used repeatedly** — Catalysts participate then regenerate, so they are not consumed — a small amount lasts.
Increasing the CONCENTRATION of reactants speeds up a reaction because:	**particles are closer together and collide more frequently** — More particles per volume means more collisions per second — a faster rate.
A catalyst does NOT change the __ of a reaction:	**products or their amount at equilibrium** — A catalyst speeds the reaction but leaves the products and equilibrium position unchanged.
On an energy diagram, the 'hump' between reactants and products represents the:	**activation energy barrier** — The peak of the curve is the activation-energy barrier the reaction must surmount.
A catalyst is NOT consumed in a reaction, meaning it:	**can be recovered unchanged afterward** — The catalyst emerges chemically unchanged, so a small amount can process a lot of reactant.
A catalyst changes an energy diagram by:	**lowering the peak (activation energy) without changing reactants or products** — The catalyzed path has a lower barrier; the starting and ending energies are unchanged.
A CATALYST speeds up a reaction by:	**providing a lower-activation-energy pathway** — A catalyst offers an alternate route with a lower energy barrier and is not consumed.
Increasing the SURFACE AREA of a solid reactant (e.g. powdering it) increases the rate because:	**more particles are exposed for collisions** — Grinding a solid exposes more surface, allowing more collisions per second.
Predicting how a change affects rate (via collision theory) lets a chemist:	**control how fast a reaction runs** — Understanding kinetics lets chemists speed up, slow down, or safely manage reactions.
Two reactions with different activation energies: the one with the LOWER activation energy will generally be:	**faster (more collisions clear the lower barrier)** — A lower barrier means more collisions have enough energy, so the reaction proceeds faster.

The activation energy is best described as the energy:	**barrier that colliding particles must overcome to react** — It is the minimum energy hill that reactant collisions must climb to proceed.
If you want to SLOW a reaction (e.g. preserve food), you might:	**lower the temperature (refrigeration)** — Cooling reduces particle energy and collision frequency — slowing spoilage reactions.
Collision theory says a reaction occurs when particles collide with:	**enough energy AND the correct orientation** — Effective collisions need sufficient energy (to break bonds) and proper orientation.
Reaction RATE measures:	**how fast reactants are converted to products over time** — Rate is the change in amount of reactant or product per unit time.
Why is kinetics essential in industry and biology?	**controlling reaction speed is critical for efficiency, safety, and life processes** — From manufacturing to metabolism, controlling how fast reactions go is central to chemistry's real-world use.
Powdering a solid reactant increases the rate by increasing the:	**surface area exposed to collisions** — Powder exposes far more surface, allowing more collisions per second.
Which factor does NOT generally increase reaction rate?	**lowering the temperature** — Lowering temperature slows particle motion and reduces effective collisions — it decreases the rate.
The rate of a reaction typically __ as the reaction proceeds and reactants are used up:	**decreases** — As reactants deplete, their concentration falls, reducing collision frequency and slowing the rate.
Temperature is a measure of the average __ of particles:	**kinetic energy** — Temperature reflects the average kinetic energy; hotter means faster-moving particles.
Only a fraction of collisions lead to reaction because many collisions:	**lack sufficient energy or have the wrong orientation** — A collision must exceed the activation energy AND be correctly oriented to react.

11. Chemical equilibrium and Le Chatelier's principle

✂ Question (fold →)	Answer
Adding a CATALYST to a system at equilibrium:	**does not shift the equilibrium (it speeds both directions equally)** — A catalyst speeds forward and reverse equally, reaching equilibrium faster but not changing its position.
The Haber process (making ammonia) uses high pressure to:	**shift equilibrium toward the fewer-gas-moles product side (more ammonia)** — $N_2 + 3H_2 \rightarrow 2NH_3$ goes from 4 gas moles to 2; high pressure favors the ammonia side.
For an ENDOthermic reaction, increasing temperature shifts equilibrium:	**toward the products (heat is a 'reactant')** — Heat is a 'reactant' for an endothermic reaction; adding it shifts the reaction forward.
Le Chatelier's principle lets a chemist:	**predict how to shift an equilibrium to favor a desired product** — By choosing conditions (concentration, temperature, pressure), chemists steer equilibria toward more product.
Predicting equilibrium shifts is powerful because it lets industry:	**maximize the yield of a valuable product** — Choosing conditions that shift equilibrium toward product maximizes yield of valuable chemicals.
A reaction at dynamic equilibrium has:	**equal forward and reverse reaction rates** — At equilibrium the forward and reverse reactions run at the same rate, so concentrations stay constant.
Increasing PRESSURE on a gaseous equilibrium shifts it toward the side with:	**fewer moles of gas** — Higher pressure favors the side with fewer gas molecules, reducing the volume pressure.
Adding more REACTANT to a system at equilibrium shifts it:	**toward the products (to consume the added reactant)** — The system shifts forward to use up the extra reactant, forming more product.
Another equilibrium misconception is that equilibrium means EQUAL concentrations. Actually:	**the rates are equal, but concentrations can be very different** — Equal RATES, not equal amounts — the equilibrium concentrations depend on the specific reaction.
Why is chemical equilibrium a central and subtle idea?	**many reactions don't go to completion but settle into a dynamic balance we can predict and shift** — Understanding the dynamic balance — and how to nudge it — is essential to controlling reversible reactions across chemistry and industry.
A common equilibrium misconception is that the reaction STOPS at equilibrium. Actually:	**both reactions continue, just at equal rates (dynamic)** — Equilibrium is dynamic: forward and reverse continue at matching rates, giving no NET change.
Removing PRODUCT from a system at equilibrium shifts it:	**toward the products (to replace what was removed)** — Removing product pulls the reaction forward to make more, replacing the loss.
A SMALL value of K ($K \ll 1$) means at equilibrium there are mostly:	**reactants** — A small K means reactants dominate — the reaction barely proceeds forward.
For an EXOthermic reaction,	

increasing temperature shifts equilibrium:	**toward the reactants (away from the added heat)** — Heat is a 'product' of an exothermic reaction; adding heat shifts it backward toward reactants.
Removing a gaseous product as it forms drives a reaction forward because it:	**keeps the system from reaching equilibrium, continually shifting toward products** — Continuously removing product keeps pulling the reaction forward (Le Chatelier), boosting yield.
The value of the equilibrium constant K depends on:	**temperature** — K changes only with temperature; catalysts and stirring don't alter its value.
Le Chatelier's principle says that if you stress a system at equilibrium, it will:	**shift to partially counteract the change** — A system at equilibrium responds to a stress by shifting in the direction that relieves it.
Equilibrium can only be reached in a:	**closed system (nothing escapes)** — In an open system products can escape, preventing a true equilibrium; a closed system is required.
At equilibrium, the concentrations of reactants and products are:	**constant over time (unchanging)** — Because forward and reverse rates match, concentrations no longer change — they're constant.
A LARGE value of K ($K \gg 1$) means at equilibrium there are mostly:	**products** — A large K means the numerator (products) dominates — the reaction favors products.
The equilibrium constant K expresses the ratio of:	**product concentrations to reactant concentrations at equilibrium** — K = [products]/[reactants] (each raised to its coefficient) at equilibrium.
At equilibrium, adding more product shifts the reaction:	**toward the reactants (to consume the added product)** — The system shifts backward to relieve the added product (Le Chatelier).
A reaction described as 'reversible' means it:	**can proceed in both the forward and reverse directions** — Reversible reactions run both ways and can settle into equilibrium.
A stress applied to an equilibrium changes the RATES temporarily until:	**a new equilibrium is established with equal rates again** — After a shift, the system settles into a new equilibrium where forward and reverse rates are equal again.
Decreasing pressure on a gaseous equilibrium shifts it toward the side with:	**more moles of gas** — Lower pressure favors the side with more gas molecules, increasing the gas volume/pressure.

12. Acids, bases, and pH

✂ Question (fold →)	Answer
A pH of 7 indicates a solution that is:	**neutral** — Pure water is neutral at pH 7 (equal H^+ and OH^-).
Pure water is neutral because it has:	**equal concentrations of H^+ and OH^-** — Water self-ionizes slightly into equal H^+ and OH^-, giving pH 7.
A STRONG acid is one that:	**ionizes completely in water** — Strength is about complete ionization (all molecules release H^+), not how much is dissolved.
The products of an acid–base neutralization are:	**a salt and water** — Neutralization yields a salt and water.
Which of these is a common property of ACIDS?	**sour taste and reaction with metals** — Acids typically taste sour, turn litmus red, and react with active metals; bases taste bitter and feel slippery.
Why are acids and bases fundamental across chemistry and biology?	**pH controls reactions in the lab, industry, the environment, and living cells** — From blood pH to soil to industrial processes, acid–base chemistry governs an enormous range of real systems.
An indicator like litmus is used to:	**signal whether a solution is acidic or basic by color** — Indicators change color depending on pH, revealing acidity or basicity.
A common pH misconception is that HIGHER pH means MORE acidic. Actually higher pH means:	**more basic (less acidic)** — The scale is inverted: lower pH = more acidic; higher pH = more basic. This is a very common error.
A BASE is a substance that, in water, produces:	**hydroxide ions (OH^-)** — An Arrhenius base releases OH^- ions in water.
A pH LESS than 7 indicates a solution that is:	**acidic** — pH below 7 means excess H^+ — acidic.
A pH GREATER than 7 indicates a solution that is:	**basic (alkaline)** — pH above 7 means excess OH^- — basic.
Bases typically feel __ and taste __:	**slippery; bitter** — Bases feel slippery and taste bitter; acids taste sour.
A titration is a technique used to:	**determine an unknown concentration by neutralizing it with a known solution** — Titration adds a measured base (or acid) until neutralization, revealing the unknown concentration.
As pH DECREASES by one unit, the H^+ concentration:	**increases tenfold** — pH is a base-10 log scale: each unit down means 10× more H^+.
A pH of 3 is how many times more acidic than a pH of 5?	**100 times (2 units × 10 each)** — Each pH unit is a factor of 10; two units means 10 × 10 = 100 times more H^+.
Pure water has a pH of:	**7 (neutral)** — Neutral pure water sits at pH 7.
When an acid and a base react, they undergo:	**neutralization, forming water and a salt** — Acid + base → water + salt is a neutralization reaction ($H^+ + OH^- \rightarrow H_2O$).
An ACID is a substance that, in water, produces:	**hydrogen ions (H^+)** — By the Arrhenius definition, an acid releases H^+ ions in water.

A common acid–base misconception is confusing STRONG with CONCENTRATED. 'Strong' means:	**the acid ionizes completely; 'concentrated' means a lot is dissolved** — A dilute strong acid and a concentrated weak acid are different — strength ≠ concentration.
A WEAK acid is one that:	**only partially ionizes in water** — A weak acid establishes an equilibrium with only some molecules ionized.
Understanding pH lets you predict outcomes because pH quantifies:	**the acidity (H^+ concentration) of a solution** — pH gives a precise, comparable measure of acidity that predicts reactivity, biology, and neutralization needs.
A solution with pH 2 is:	**strongly acidic** — pH 2 is far below 7 — strongly acidic.
The pH scale typically runs from:	**0 to 14** — The common pH scale runs 0–14, with 7 neutral.
The reaction $H^+ + OH^- \rightarrow H_2O$ represents the essence of:	**neutralization** — The net ionic equation for neutralization is H^+ combining with OH^- to form water.
Antacids relieve heartburn because they are __ that neutralize stomach acid:	**bases** — Antacids are mild bases that neutralize excess stomach H^+.

13. Thermochemistry — energy changes in reactions

✂ Question (fold →)	Answer
The activation energy and the ΔH of a reaction are:	**different quantities (barrier height vs net energy change)** — Activation energy is the barrier to react; ΔH is the net energy difference between products and reactants.
A reaction that makes the surroundings feel COLD is:	**endothermic (it absorbs heat)** — Absorbing heat from the surroundings makes them feel cold — endothermic.
The Law of Conservation of Energy says energy in a reaction is:	**transformed, not created or destroyed** — Energy changes form (chemical ↔ heat) but the total is conserved.
On an energy diagram, an ENDOthermic reaction has products at a __ energy than reactants:	**higher** — Endothermic products sit higher (energy was absorbed into them).
Photosynthesis, which stores energy from sunlight in glucose, is:	**endothermic (it absorbs energy)** — Photosynthesis absorbs light energy to build glucose — an endothermic (energy-storing) process.
A reaction with a positive ΔH will make its surroundings feel:	**colder (it absorbs heat)** — Positive ΔH is endothermic — it pulls heat from the surroundings, which feel colder.
Whether a reaction absorbs or releases energy depends on comparing:	**the energy to break reactant bonds vs the energy released forming product bonds** — The net energy change is the balance between energy invested to break bonds and energy recovered forming new ones.
A NEGATIVE ΔH corresponds to a reaction that is:	**exothermic (releases energy)** — By convention, releasing energy gives a negative ΔH (energy left the system).
An EXOthermic reaction:	**releases energy (usually as heat) to the surroundings** — Exothermic reactions give off energy, so the surroundings get warmer.
Predicting whether a reaction is endo- or exothermic lets a chemist:	**anticipate heating/cooling and energy requirements** — Knowing the energy direction guides safety, reactor design, and process energy needs.
The energy stored in the bonds of reactants is a form of:	**chemical potential energy** — Bonds hold chemical potential energy, released or absorbed when they rearrange.
On an energy diagram, an EXOthermic reaction has products at a __ energy than reactants:	**lower** — Exothermic products sit lower on the energy diagram (energy was released).
A reaction that makes the surroundings feel HOT is:	**exothermic (it releases heat)** — Releasing heat warms the surroundings — exothermic.
A reaction is EXOthermic overall when:	**more energy is released forming bonds than is used breaking them** — If bond-forming releases more than bond-breaking absorbs, the net reaction gives off energy.
A cold pack works by using a reaction (or dissolving) that is:	**endothermic (absorbs heat, feeling cold)** — A cold pack uses an endothermic process that absorbs heat from its surroundings, feeling cold.
Why is thermochemistry essential across science and technology?	**energy changes power engines, bodies, and industry, and must be understood and controlled** — Every fuel, battery, metabolic process, and industrial reaction is governed by its energy changes.

A hand warmer works by using a reaction that is:	**exothermic (releases heat, feeling warm)** — A hand warmer uses an exothermic reaction that releases heat.
If a reaction beaker gets warmer during a reaction, the reaction is:	**exothermic** — Heat released into the beaker (warming it) means the reaction is exothermic.
Energy is conserved in every reaction, so the energy released or absorbed equals the:	**difference between reactant and product bond energies** — The net energy change is the balance of energy to break versus form bonds.
A common thermochemistry misconception is that exothermic reactions 'create' energy. Actually they:	**convert stored chemical energy into heat (energy is conserved)** — The heat released was stored in the reactants' bonds — energy is converted, not created.
A POSITIVE ΔH corresponds to a reaction that is:	**endothermic (absorbs energy)** — Absorbing energy gives a positive ΔH (energy entered the system).
Breaking chemical bonds __ energy, and forming bonds __ energy:	**requires (absorbs); releases** — Breaking bonds costs energy (input); forming bonds gives energy back (output).
The energy absorbed or released by a reaction is called the:	**enthalpy change (ΔH)** — ΔH (enthalpy change) is the heat absorbed or released at constant pressure.
An ENDOthermic reaction:	**absorbs energy from the surroundings** — Endothermic reactions take in energy, so the surroundings get cooler.
Combustion (burning fuel) is an example of a reaction that is:	**exothermic** — Burning fuel releases large amounts of heat — strongly exothermic.

14. Redox — oxidation, reduction, and electron transfer

✂ Question (fold →)	Answer
An element whose oxidation number DECREASES during a reaction has been:	**reduced (it gained electrons)** — A drop in oxidation number means the element gained electrons — reduction.
Corrosion (like rusting) is an unwanted redox process that:	**oxidizes and degrades a metal over time** — Corrosion is the oxidation of a metal by its environment, weakening it.
A redox reaction always involves:	**the transfer of electrons from one species to another** — Redox = reduction + oxidation; electrons move from the substance oxidized to the one reduced.
The oxidation number of oxygen in most compounds is:	**−2** — Oxygen is usually −2 in compounds (with a few exceptions like peroxides).
When iron rusts ($Fe \rightarrow Fe^{3+}$), the iron is:	**oxidized (it loses electrons)** — Going from Fe^0 to Fe^{3+} loses three electrons — oxidation.
In OIL RIG, the 'RIG' stands for:	**Reduction Is Gain (of electrons)** — RIG = Reduction Is Gain of electrons (OIL = Oxidation Is Loss).
The substance that GAINS electrons (is reduced) is called the:	**oxidizing agent** — The oxidizing agent causes oxidation by taking electrons — so it is itself reduced.
To track electron transfer in a reaction, chemists assign:	**oxidation numbers to each element** — Changes in oxidation numbers reveal which species is oxidized and which is reduced.
A reducing agent is itself __ during the reaction:	**oxidized** — The reducing agent donates electrons, so it is itself oxidized.
The oxidation number of hydrogen in most compounds is:	**+1** — Hydrogen is usually +1 in compounds (−1 only with metals in hydrides).
In a battery/electrochemical cell, oxidation occurs at the __ and reduction at the __:	**anode; cathode** — Oxidation happens at the anode; reduction happens at the cathode ('an ox, red cat').
A battery generates electricity by using a redox reaction to:	**push electrons through an external circuit** — In a battery, oxidation at one electrode releases electrons that flow through the circuit to reduction at the other.
If one substance is oxidized in a reaction, another MUST be:	**reduced (it gains the transferred electrons)** — Electrons lost by one species are gained by another, so oxidation and reduction always occur together.
An increase in an element's oxidation number indicates it was:	**oxidized** — A rise in oxidation number means the element lost electrons — oxidation.
Electroplating uses redox to:	**deposit a metal coating onto an object by reduction** — In electroplating, metal ions are reduced onto the object's surface, forming a coating.
Why is redox one of the most important reaction classes in	**electron transfer underlies energy production, corrosion, metabolism, and electrochemistry** — From the batteries in your devices to the metabolism in

chemistry?	your cells, redox electron transfer is everywhere energy moves.
The oxidation number of an element in its FREE (uncombined) state is:	**0** — A pure element (like O_2 or Fe metal) has an oxidation number of 0.
Redox reactions are essential to life and technology because they:	**transfer energy via electron flow (respiration, batteries, photosynthesis)** — Electron transfer powers cellular respiration, photosynthesis, batteries, and metal production.
The substance that LOSES electrons (is oxidized) is called the:	**reducing agent** — The reducing agent donates electrons — so it is itself oxidized.
REDUCTION is defined as the:	**gain of electrons** — Reduction is the gain of electrons (remember 'RIG': Reduction Is Gain).
In the reaction Zn + Cu^{2+} → Zn^{2+} + Cu, the zinc is:	**oxidized (Zn → Zn^{2+})** — Zinc goes from 0 to +2, losing electrons — it is oxidized (and is the reducing agent).
The overall electron accounting in a redox reaction must:	**balance — electrons lost equal electrons gained** — Every electron lost by the oxidized species is gained by the reduced one — they balance.
In that same reaction, the copper ion Cu^{2+} is:	**reduced (Cu^{2+} → Cu)** — Cu^{2+} gains two electrons to become Cu metal — reduction.
OXIDATION is defined as the:	**loss of electrons** — Oxidation is the loss of electrons (remember 'OIL': Oxidation Is Loss).
A common redox misconception is that oxidation always involves oxygen. Actually oxidation:	**is any loss of electrons, whether or not oxygen is involved** — The modern definition is electron loss; the historical 'gaining oxygen' is just one example.

15. Introductory organic chemistry — carbon and functional groups

✂ Question (fold →)	Answer
Organic chemistry is the study of compounds containing:	**carbon** — Organic chemistry centers on carbon-containing compounds (with hydrogen and other elements).
Functional groups let chemists predict a molecule's reactivity because:	**the same functional group behaves similarly across different molecules** — Recognizing a functional group tells you how a molecule will react, regardless of the rest of its structure.
The simplest hydrocarbon, methane, has the formula:	**CH_4** — Methane is one carbon bonded to four hydrogens: CH_4.
Plastics, proteins, and DNA are all examples of:	**polymers (large repeating-unit molecules)** — These are all polymers — long chains of repeating monomer units.
A compound with the formula C_2H_6 (only single bonds) is an:	**alkane (ethane)** — Only single C–C bonds and the -ane pattern make C_2H_6 an alkane (ethane).
An ALKENE contains at least one:	**carbon–carbon double bond** — Alkenes have a C=C double bond (e.g. ethene), making them unsaturated.
The functional group that makes a molecule an amine contains:	**nitrogen** — Amines contain a nitrogen-based functional group (–NH_2).
An ALKANE is a hydrocarbon with:	**only single carbon–carbon bonds** — Alkanes contain only single C–C bonds (saturated hydrocarbons), like methane and ethane.
The –COOH functional group defines the family of:	**carboxylic acids** — The carboxyl group (–COOH) makes a molecule a carboxylic acid (like acetic acid).
Ethanol (C_2H_5OH) belongs to which family?	**alcohols (it has an –OH group)** — The –OH group makes ethanol an alcohol.
A polymer is a large molecule made of:	**many repeating smaller units (monomers)** — Polymers (like plastics and proteins) are long chains of repeating monomer units.
Combustion of a hydrocarbon fuel (like octane in gasoline) produces:	**carbon dioxide, water, and energy** — Burning a hydrocarbon fuel in oxygen yields CO_2, H_2O, and released energy.
A FUNCTIONAL GROUP is:	**a specific atom grouping that gives a molecule characteristic properties** — Functional groups (like –OH, –COOH) are reactive atom clusters that define a molecule's chemistry.
Carbon's ability to form long chains is called:	**catenation** — Catenation is carbon's ability to bond to itself in chains and rings — the basis of organic diversity.
Carbon is uniquely versatile in forming molecules because it can form:	**four covalent bonds, including with other carbons (chains and rings)** — Carbon's four bonds and ability to bond to itself let it build enormous chains, branches, and rings.
An ALKYNE contains at least one:	**carbon–carbon triple bond** — Alkynes have a C≡C triple bond (e.g. ethyne/acetylene).
A common organic misconception is	**carbon-containing (whether natural or synthetic)** — In chemistry,

that 'organic' means 'natural/healthy.' In chemistry, organic means:	'organic' simply means carbon-based; many organic compounds are synthetic (plastics, medicines).
Isomers matter because molecules with the same formula can have:	**very different properties and uses** — Arrangement changes behavior — isomers can differ dramatically in properties despite sharing a formula.
Two molecules that are isomers will generally have:	**the same molecular formula but different properties** — Isomers share a formula but differ in structure, giving different properties.
ISOMERS are molecules with:	**the same formula but different structures** — Isomers share a molecular formula but differ in how atoms are arranged, giving different properties.
The –OH functional group defines the family of:	**alcohols** — The hydroxyl group (–OH) makes a molecule an alcohol (like ethanol).
A HYDROCARBON contains only:	**carbon and hydrogen** — Hydrocarbons are made solely of carbon and hydrogen (e.g. methane, octane).
Saturated hydrocarbons are 'saturated' because they contain:	**the maximum number of hydrogen atoms (all single bonds)** — 'Saturated' means every carbon holds as many hydrogens as possible — only single bonds.
Why is organic chemistry central to life and industry?	**carbon compounds make up living things, fuels, plastics, and medicines** — The chemistry of carbon underlies all life, our fuels, materials, and pharmaceuticals — an enormous field.
The huge diversity of organic compounds arises because carbon can form:	**chains, branches, and rings of nearly unlimited size** — Carbon's four self-linking bonds allow an essentially unlimited variety of stable structures.

16. Capstone — integrating stoichiometry, energy, and equilibrium

✂ Question (fold →)	Answer
The 'predict–observe–explain' habit in the lab means you:	**predict the result, run the reaction, then reconcile any surprise** — Predicting before running — then explaining discrepancies — turns a procedure into genuine understanding.
The ultimate goal of chemistry as a science is to:	**understand, predict, and control the transformations of matter** — Chemistry seeks to understand matter's behavior well enough to predict and control chemical change.
A key safety-and-reasoning step is to PREDICT the products before mixing reactants so you can:	**anticipate hazards and check the observed result against expectation** — Predicting products supports both safety and the predict–observe–explain learning loop.
Johnstone's triangle reminds chemists to connect three levels:	**macroscopic (what you see), submicroscopic (particles), and symbolic (formulas)** — Deep understanding links the observed phenomenon, the particle-level explanation, and the symbolic representation.
Before running a reaction, converting all given masses to moles lets you:	**compare amounts using the balanced mole ratio** — Moles are the currency of the balanced equation, so convert to moles before comparing.
A reaction that is exothermic AND increases disorder tends to be:	**spontaneous (favorable)** — Releasing energy and increasing disorder both favor a reaction proceeding on its own.
Integrating stoichiometry with kinetics: to make a needed amount of product FASTER, you can:	**raise temperature or add a catalyst while keeping the correct mole ratios** — Stoichiometry sets the amounts; kinetics sets the speed — adjust conditions without breaking the ratios.
Before scaling a reaction up, a chemist checks the __ to know reactant proportions:	**balanced equation's mole ratios** — Mole ratios from the balanced equation set the correct proportions at any scale.
If a gas is produced and escapes an OPEN container, the measured mass appears to:	**decrease, though mass is truly conserved** — The escaping gas carries away mass; in a sealed system the total would be unchanged.
Running a real lab reaction integrates skills: balance, find the limiting reagent, predict yield, and:	**account for energy and equilibrium effects** — A complete analysis combines stoichiometry, energetics, and equilibrium — the whole toolkit at once.
Reasoning responsibly about a chemical model means acknowledging that predictions:	**assume ideal conditions and may differ from messy real results** — Good chemistry states the assumptions (ideal gas, complete reaction) and expects real results to vary.
Combining kinetics and	**reaching equilibrium FASTER without changing the equilibrium position** — A

equilibrium: a catalyst helps a reversible reaction by:	catalyst speeds both directions equally — you reach the same equilibrium sooner, not a different one.
Integrating concepts, to MAXIMIZE product in a reversible exothermic reaction you might:	**use lower temperature (favors products) but a catalyst to keep the rate practical** — Lower temperature shifts an exothermic equilibrium toward products, but a catalyst keeps the rate workable — a real engineering trade-off.
The FIRST step before doing any quantitative chemistry on a reaction is to:	**write and balance the chemical equation** — A balanced equation supplies the mole ratios and conserves mass — the foundation for all calculations.
Checking that your predicted product amount is reasonable involves confirming the:	**units and that mass is conserved** — Verifying units and mass conservation catches setup errors — part of reasoning about, not just computing, a result.
To predict how conditions shift a reversible reaction's yield, you apply:	**Le Chatelier's principle** — Le Chatelier predicts how changing concentration, temperature, or pressure shifts the equilibrium yield.
If a real reaction gives a 70% yield, the missing 30% is explained by:	**side reactions, incomplete reaction, or losses in handling** — Real yields fall short due to competing reactions, equilibrium limits, and recovery losses — not lost atoms.
Why does chemistry let us predict and control the material world so powerfully?	**because matter follows conserved, quantifiable rules — atoms, energy, and equilibrium** — Conservation laws, stoichiometry, energetics, and equilibrium make chemical change predictable and controllable — the power of chemistry.
A safety-first simulated lab lets students:	**explore reactions and predictions without physical hazards** — A simulation allows safe, repeatable exploration of predictions and outcomes — the POE loop without hazard.
The biggest reasoning habit ReactQuest builds is to:	**predict what a reaction will do, then run it and reconcile the result** — Predict-then-reconcile is the scientific habit that turns lab procedures into genuine chemical understanding.
Conservation of mass lets you check a reaction result by verifying that:	**total product mass equals total reactant mass (in a closed system)** — Because atoms are conserved, the masses must balance — a useful sanity check on any result.
To predict whether a reaction will release or absorb heat, you consider its:	**enthalpy change (endothermic vs exothermic)** — The sign of ΔH tells you whether the reaction releases (exothermic) or absorbs (endothermic) energy.
The mole is the central bridge in lab chemistry because it connects:	**measurable mass to countable particles and reaction ratios** — The mole links what you can weigh (grams) to what reacts (particles and ratios) — the linchpin of quantitative chemistry.
To predict how much product a reaction can make, you use:	**stoichiometry from the limiting reagent** — Stoichiometry starting from the limiting reagent gives the theoretical yield.
To predict how FAST a reaction proceeds, you consider:	**kinetics (temperature, concentration, catalyst, surface area)** — Kinetics — via collision theory — predicts the reaction speed from the conditions.

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