FOUNDATIONS OF CHEMISTRY

A High School Introduction to Chemical Science — with Bridges to Physics

2026

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How to Use This Book

This course is designed for high school students encountering chemistry for the first time. Every chapter builds directly on the one before it, so working in order matters. Each chapter includes:

- **Plain-language explanations** of every concept before any formulas appear

- **Worked examples** showing every step, including the thinking behind each step
- **Physics Bridges** at the end of each chapter — connections between what you just learned in chemistry and what you will encounter in physics
- **Review Questions** to test understanding before moving on

No prior science background is assumed. You need arithmetic and basic algebra (solving for one variable). Everything else is introduced as you need it.

A note on units: chemistry and physics both depend on precise measurement. This course uses the **International System of Units (SI)** — meters, kilograms, seconds, moles, and their derivatives — because that is the language of every scientific field worldwide.

Chapter One: What Is Chemistry — and Why Does It Matter?

1.1 The Central Question

Chemistry is the scientific study of **matter** — what it is made of, how it behaves, and how it changes. That definition sounds simple, but it covers almost everything in the physical universe: the air you breathe, the food your cells burn for energy, the steel in bridges, the silicon in your phone, the water in your blood, and the carbon in every living thing on Earth.

The reason chemistry is taught before physics in most curricula is straightforward: physics describes how objects move, how forces act, how energy flows — but all of those objects are made of matter. Before you can understand why a rocket engine works, you need to understand what happens when fuel burns. Before you can understand electricity in a circuit, you need to understand why some materials conduct electrons and others do not. Chemistry provides the foundation.

1.2 Matter: The Stuff of Everything

Matter is anything that has mass and takes up space (volume). A rock is matter. Water is matter. The air in a balloon is matter. Light is *not* matter — it has no mass and takes up no space. Gravity is not matter. These distinctions will become important when we reach energy.

States of Matter

Matter exists in four states, though three are encountered in everyday chemistry:

Solid: Particles are packed tightly together in a fixed arrangement. A solid has a definite shape and a definite volume. Ice is a solid.

Liquid: Particles are close together but can flow past one another. A liquid has a definite volume but takes the shape of its container. Water is a liquid.

Gas: Particles are widely separated and move freely at high speeds. A gas has neither a definite shape nor a definite volume — it expands to fill whatever container holds it. Water vapor (steam) is a gas.

Plasma: A high-energy state in which electrons are stripped from atoms. Plasma is the most common state of matter in the universe (stars are plasma) but the rarest in everyday life on Earth.

Example 1.1 — Identifying States of Matter

Identify the state of matter: (a) a gold coin; (b) gasoline in a tank; (c) the oxygen you exhale; (d) the flame of a candle.

(a) Gold coin: Solid — definite shape, definite volume, particles locked in place.

(b) Gasoline: Liquid — definite volume, flows to fill the bottom of its container.

(c) Exhaled oxygen (and carbon dioxide): Gas — no definite shape or volume, expands freely.

(d) Candle flame: Plasma — hot ionized gas. (Most people say "gas," which is close enough at this level.)

1.3 Physical vs. Chemical Changes

Not all changes to matter are the same kind. Chemistry distinguishes two fundamental types:

Physical change: The form or appearance of matter changes, but its chemical identity does not. Ice melting into water is a physical change — you still have H_2O, just in a different state. Cutting a piece of paper is a physical change. Dissolving sugar in water is a physical change (the sugar is still sugar; it just disperses).

Chemical change (chemical reaction): The chemical identity of matter changes — new substances form. When wood burns, the cellulose in the wood combines with oxygen and produces carbon dioxide, water vapor, and ash. The wood is gone; you cannot "un-burn" it. Rusting iron, digesting food, baking a cake, and an explosion are all chemical changes.

Signs that a Chemical Change Has Occurred

Because we cannot always see individual atoms rearranging, chemists look for observable evidence:

- A new color appears
- A gas is produced (bubbling when no heat was applied)
- A solid forms in a liquid (called a **precipitate**)
- Heat or light is released or absorbed in unexpected amounts
- The change is difficult or impossible to reverse

Example 1.2 — Physical or Chemical Change?

(a) Crushing an aluminum can; (b) burning magnesium ribbon in air; (c) dissolving salt in water; (d) milk souring.

(a) Crushing a can: Physical — aluminum is still aluminum; only the shape changes.

(b) Burning magnesium: Chemical — magnesium reacts with oxygen to form magnesium oxide (MgO), a white powder entirely different from the shiny metal you started with.

(c) Dissolving salt: Physical — the sodium chloride disperses into ions in solution, but no new substance forms. Evaporate the water and the salt returns.

(d) Milk souring: Chemical — bacteria convert lactose (a sugar) into lactic acid. The milk's chemical composition has fundamentally changed.

1.4 Pure Substances vs. Mixtures

Pure substance: Has a uniform and definite chemical composition throughout. Every sample of pure water is exactly H_2O. Every sample of pure gold is exactly Au. Pure substances are either elements or compounds (more on these in Chapter 2).

Mixture: Two or more substances combined without a fixed ratio and without a chemical reaction. Mixtures can be separated by physical means.

Mixtures are further divided:

Homogeneous mixture (solution): Uniform throughout at the molecular scale. Saltwater, air, and brass (a mixture of copper and zinc) are homogeneous. You cannot see the individual components.

Heterogeneous mixture: Components are visibly distinct. Sand in water, granite, and a bowl of cereal are heterogeneous.

Example 1.3 — Classifying Matter

Classify each: (a) pure copper wire; (b) a glass of lemonade; (c) a salad; (d) carbon dioxide gas.

(a) Copper wire: Pure substance (an element).

(b) Lemonade: Homogeneous mixture — water, sugar, lemon juice, and other components are uniformly distributed.

(c) A salad: Heterogeneous mixture — you can see the individual lettuce, tomatoes, croutons, etc.

(d) Carbon dioxide: Pure substance (a compound — two oxygen atoms bonded to one carbon atom).

1.5 Measurement and Scientific Notation

Chemistry is a quantitative science. Every experiment produces measurements. The most important measuring skill in introductory chemistry is using **SI units** and expressing very large or very small numbers in **scientific notation**.

SI Units Relevant to Chemistry

Quantity	SI Unit	Symbol
Length	meter	m
Mass	kilogram	kg

Quantity	SI Unit	Symbol
Temperature	kelvin	K
Amount of substance	mole	mol
Volume	liter (derived)	L
Energy	joule	J

Scientific Notation

A number in scientific notation is written as:

$A \times 10^{n}$ where **$1 \leq A < 10$** and **n** is an integer.

Example 1.4 — Scientific Notation

(a) Write 602,200,000,000,000,000,000,000 in scientific notation. (b) Write 0.000000000529 in scientific notation.

(a) Count the places the decimal moves left until one digit remains before the decimal point:

602,200,000,000,000,000,000,000 → **6.022×10^{23}**

This number — 6.022×10^{23} — is **Avogadro's number**, one of the most important constants in chemistry.

(b) Count the places the decimal moves right:

0.000000000529 → **5.29×10^{-10}**

This is the radius of a hydrogen atom in meters.

1.6 Significant Figures

Every measurement has some uncertainty. **Significant figures** (sig figs) communicate how precise a measurement is. The rules:

1. All non-zero digits are significant: 4.53 has 3 sig figs.
2. Zeros between non-zero digits are significant: 4,053 has 4 sig figs.
3. Leading zeros are never significant: 0.0045 has 2 sig figs.
4. Trailing zeros after a decimal point are significant: 4.530 has 4 sig figs.

Arithmetic rules: - *Multiplication/division:* The answer has the same number of sig figs as the measurement with the fewest sig figs. - *Addition/subtraction:* The answer is rounded to the same decimal place as the least precise measurement.

Example 1.5 — Significant Figures

Calculate: (a) 4.52 × 2.1; (b) 12.356 + 0.07

(a) 4.52 has 3 sig figs; 2.1 has 2 sig figs. Answer must have 2 sig figs. 4.52 × 2.1 = 9.492 → rounded to **9.5**

(b) 12.356 goes to the thousandths place; 0.07 goes to the hundredths place. Round to hundredths. 12.356 + 0.07 = 12.426 → rounded to **12.43**

Physics Bridge 1: Chemistry and Physics Share the Same Foundation

Physics and chemistry both study matter and energy. The distinction between them is mostly one of emphasis:

- **Chemistry** focuses on the *composition* of matter and the *transformations* between substances.
- **Physics** focuses on the *behavior* of matter — its motion, forces, and energy transfers.

The states of matter you learned here (solid, liquid, gas) reappear in physics in the study of **thermodynamics** — the science of heat and energy. When you study the ideal gas law in physics, you will use the same understanding of gases developed in chemistry. The concept of **energy** introduced as "heat released or absorbed in a chemical reaction" becomes a central organizing principle in all of physics.

Scientific notation and significant figures are tools both disciplines share. In physics, you will encounter numbers as large as the speed of light (3.00×10^{8} m/s) and as small as the mass of a proton (1.67×10^{-27} kg). The skills you built in this chapter are the precision tools of all physical science.

Chapter 1 Review Questions

1. Define matter. Give two examples of something that is matter and two that are not.
2. A sample of iron rusts when exposed to moist air. Is this a physical or chemical change? Identify two pieces of observable evidence that support your answer.
3. Classify each as a pure substance or mixture; if a mixture, state whether it is homogeneous or heterogeneous: (a) tap water; (b) diamond; (c) concrete; (d) table sugar.
4. Write 0.00000472 in scientific notation.
5. How many significant figures are in each: (a) 0.0430; (b) 1,050; (c) 7.00×10^{4}?
6. Calculate 3.14 × 2.5 and express the answer with the correct number of significant figures.

Chapter Two: Atoms, Elements, and the Periodic Table

2.1 The Atomic Theory — A Brief History

The idea that matter is made of tiny, indivisible particles is ancient — the Greek philosopher Democritus proposed it around 400 B.C.E. — but it remained a philosophical idea until John Dalton gave it a scientific foundation in 1803.

Dalton's Atomic Theory (simplified): 1. All matter is made of atoms — tiny, indivisible particles. 2. All atoms of the same element are identical (same mass, same properties). 3. Atoms of different elements are different. 4. Chemical reactions rearrange atoms; no atoms are created or destroyed. 5. Compounds form when atoms of different elements combine in fixed, whole-number ratios.

Most of Dalton's theory survives today. The major revision: atoms are *not* indivisible — they have internal structure. But point 4 (conservation of atoms in chemical reactions) is as true today as it was in 1803.

2.2 Inside the Atom: Protons, Neutrons, and Electrons

By the early twentieth century, experiments revealed that atoms contain three types of particles:

Particle	Symbol	Charge	Relative Mass	Location
Proton	p^+	+1	1 amu	Nucleus
Neutron	n^0	0	1 amu	Nucleus
Electron	e^-	−1	1/1836 amu	Electron cloud

amu = atomic mass unit. One amu is defined as exactly 1/12 the mass of a carbon-12 atom.

The **nucleus** is the dense, positively charged core of the atom. It contains protons and neutrons and is incredibly small — if an atom were the size of a football stadium, the nucleus would be a marble at the center. Nearly all of the atom's mass is concentrated in this tiny nucleus.

The **electron cloud** surrounds the nucleus. Electrons are so light that they contribute almost nothing to atomic mass, but they determine *all* of an atom's chemical behavior.

2.3 Atomic Number and Mass Number

Atomic number (Z): The number of protons in an atom's nucleus. This number *defines* the element — every atom of hydrogen has exactly 1 proton; every atom of oxygen has exactly 8; every atom of gold has exactly 79.

Mass number (A): The total number of protons *plus* neutrons in the nucleus.

A = protons + neutrons

Isotopes: Atoms of the same element (same number of protons) with different numbers of neutrons. Carbon-12 has 6 protons and 6 neutrons (mass number = 12). Carbon-14 has 6 protons

and 8 neutrons (mass number = 14). Both are carbon; they have identical chemical properties but different masses.

Example 2.1 — Protons, Neutrons, Electrons

An atom of phosphorus has the symbol $^{31}_{15}P$. How many protons, neutrons, and electrons does a neutral atom have?

- Atomic number (subscript) = **15** → 15 protons
- Mass number (superscript) = **31** → 31 − 15 = **16 neutrons**
- Neutral atom: electrons = protons = **15 electrons**

Example 2.2 — Isotopes

Chlorine has two naturally occurring isotopes: Cl-35 and Cl-37. How many neutrons does each have?

- Cl-35: Z = 17, A = 35, neutrons = 35 − 17 = **18**
- Cl-37: Z = 17, A = 37, neutrons = 37 − 17 = **20**

Both atoms have 17 protons and behave identically in chemical reactions.

2.4 Atomic Mass

The **atomic mass** shown on the periodic table is not a whole number because it is a *weighted average* of the masses of all naturally occurring isotopes, weighted by their natural abundance.

Example 2.3 — Calculating Atomic Mass

Chlorine is 75.77% Cl-35 (mass = 34.969 amu) and 24.23% Cl-37 (mass = 36.966 amu). Calculate the atomic mass of chlorine.

Convert percentages to decimals, then multiply each isotope's mass by its fractional abundance:

Atomic mass = (0.7577 × 34.969) + (0.2423 × 36.966)

= 26.496 + 8.956

= **35.45 amu**

This matches the value on the periodic table.

2.5 The Periodic Table — Organization and Patterns

The **periodic table** arranges all known elements (118 as of 2026) by increasing atomic number, with elements having similar chemical properties placed in the same vertical column.

Reading the Periodic Table

Each box contains: - The element's **symbol** (1–2 letters) - The **atomic number** (upper number) - The **atomic mass** (lower number, in amu)

Periods and Groups

Periods (rows): Elements in the same period have the same number of electron **shells**. Period 1 has 2 elements; Period 2 has 8; Period 3 has 8.

Groups (columns): Elements in the same group have similar chemical properties because they have the same number of **valence electrons** (outermost electrons — the ones that participate in bonding). Groups are numbered 1–18.

Key Groups and Their Names

Group	Common Name	Characteristics
1	Alkali Metals	Soft, very reactive, form +1 ions
2	Alkaline Earth Metals	Harder, reactive, form +2 ions
17	Halogens	Very reactive nonmetals, form −1 ions
18	Noble Gases	Colorless, odorless, nearly unreactive
3–12	Transition Metals	Include iron, copper, gold, silver

Metals, Nonmetals, and Metalloids

The periodic table is divided by a staircase line running from boron (B) to astatine (At):

- **Metals** (left and center): Shiny, conduct electricity and heat, malleable, ductile. Most elements are metals.
- **Nonmetals** (upper right): Poor conductors, brittle as solids. Include oxygen, nitrogen, carbon, and the halogens.
- **Metalloids** (along the staircase: B, Si, Ge, As, Sb, Te): Intermediate properties; silicon is the basis of semiconductor technology.

Example 2.4 — Using the Periodic Table

For the element selenium (Se): (a) What is its atomic number? (b) What period and group? (c) Is it a metal, nonmetal, or metalloid? (d) How many protons and electrons does a neutral atom have?

Selenium: atomic number = **34**, located in **Period 4, Group 16**. Selenium is along the metalloid staircase region — it is classified as a **nonmetal** (though it has some metalloid character). Neutral selenium: **34 protons, 34 electrons**.

2.6 Electron Configuration — Where Electrons Live

Electrons occupy regions of space called **orbitals**, grouped into **shells** (energy levels) numbered 1, 2, 3, 4… outward from the nucleus. Each shell can hold a maximum number of electrons:

- Shell 1: maximum **2** electrons

- Shell 2: maximum **8** electrons
- Shell 3: maximum **8** electrons (for elements in periods 1–3)
- Shell 4 and beyond: up to **18** or more

The electrons in the **outermost shell** are the **valence electrons**. They determine how an element bonds.

Example 2.5 — Electron Configuration by Shell

Write the shell-by-shell electron configuration for sodium (Na, Z = 11) and chlorine (Cl, Z = 17).

Sodium (11 electrons): - Shell 1: 2 - Shell 2: 8 - Shell 3: 1 Configuration: **2, 8, 1** → 1 valence electron

Chlorine (17 electrons): - Shell 1: 2 - Shell 2: 8 - Shell 3: 7 Configuration: **2, 8, 7** → 7 valence electrons

This immediately explains why sodium (1 valence e^-) readily gives its electron to chlorine (which needs 1 more to complete its outer shell) — they react vigorously to form sodium chloride (table salt, NaCl).

Physics Bridge 2: Atoms in Physics

In physics, the atom you learned here is the foundation of several major topics:

Nuclear physics deals with the nucleus — the protons and neutrons you just studied. Nuclear reactions (fission, fusion) involve changes to the nucleus itself, releasing enormous energy described by Einstein's famous equation $E = mc^2$.

Electricity is the study of electrons in motion. The electrical conductivity of metals comes directly from their electron configurations — metals have loosely held valence electrons that can move freely through the material.

Quantum mechanics — the branch of physics that replaces classical physics at the atomic scale — was developed precisely because classical physics could not explain the behavior of electrons in atoms. The electron shells you learned here are a simplified picture of quantum mechanical orbitals.

Chapter 2 Review Questions

1. State the four key points of Dalton's atomic theory that are still accepted today.
2. What distinguishes one element from another at the atomic level?
3. An atom has 26 protons and 30 neutrons. What element is it? What is its mass number? Write the isotope symbol.

4. Copper has two isotopes: Cu-63 (abundance 69.17%, mass 62.93 amu) and Cu-65 (abundance 30.83%, mass 64.93 amu). Calculate the atomic mass of copper.

5. Without looking at a periodic table, predict whether sodium (Na) or chlorine (Cl) is more likely to lose an electron in a chemical reaction. Explain using valence electrons.

6. Identify the group number, period number, and metal/nonmetal/metalloid classification for: (a) Calcium; (b) Neon; (c) Silicon; (d) Bromine.

Chapter Three: Chemical Bonding — How Atoms Join Together

3.1 Why Atoms Bond

Isolated atoms are often chemically unstable. The key to stability, for most atoms, is achieving a **full outer electron shell**. The noble gases (Group 18) already have full outer shells — helium has 2 valence electrons (a full shell 1), and neon, argon, krypton, xenon, and radon all have 8 valence electrons. This explains why noble gases are nearly chemically inert.

Other atoms achieve this stability by **gaining, losing, or sharing electrons** with other atoms. This electron redistribution is what we call a **chemical bond**.

The **octet rule** states that most atoms tend to form bonds until they are surrounded by 8 valence electrons (an octet). Hydrogen is the most important exception — it needs only 2 valence electrons to fill its first shell.

3.2 Ionic Bonds — Electron Transfer

An **ionic bond** forms when one atom *transfers* one or more electrons to another atom, creating oppositely charged particles that attract each other.

- The atom that *loses* electrons becomes a positively charged **cation** (ca-TIGH-on).
- The atom that *gains* electrons becomes a negatively charged **anion** (AN-eye-on).

Metals tend to form cations (they have few valence electrons and lose them easily).
Nonmetals tend to form anions (they have nearly full shells and gain electrons easily).

Example 3.1 — Forming an Ionic Bond

Explain, step by step, how sodium and chlorine form sodium chloride (NaCl).

Sodium (Na): Group 1, configuration 2, 8, 1 → has 1 valence electron. Losing it gives the stable configuration 2, 8 (same as neon).

$Na \rightarrow Na^+ + e^-$

Chlorine (Cl): Group 17, configuration 2, 8, 7 → needs 1 more electron to reach 2, 8, 8 (same as argon).

$Cl + e^- \rightarrow Cl^-$

Electron transfer: Sodium donates its electron to chlorine. Result: Na^+ and Cl^-.

Electrostatic attraction: Opposite charges attract. Na^+ and Cl^- pull toward each other and are held together in a crystal lattice — the structure of table salt.

Ionic compounds form crystalline solids at room temperature and have high melting points. They conduct electricity when dissolved in water or melted (because the ions become free to move).

Example 3.2 — Predicting Ionic Charge

Predict the most common ionic charge for: (a) calcium (Group 2); (b) oxygen (Group 16); (c) aluminum (Group 13).

(a) Calcium: 2 valence electrons → loses 2 → **Ca^{2+}**

(b) Oxygen: 6 valence electrons → gains 2 to reach 8 → **O^{2-}**

(c) Aluminum: 3 valence electrons → loses 3 → **Al^{3+}**

3.3 Covalent Bonds — Electron Sharing

A **covalent bond** forms when two atoms *share* electrons. Covalent bonding occurs primarily between nonmetal atoms. Rather than giving electrons away, each atom contributes one or more electrons to a shared pair that "belongs" to both atoms simultaneously.

Single bond: 1 shared pair (2 electrons). Written as a single line: H–H

Double bond: 2 shared pairs (4 electrons). Written as a double line: O=O

Triple bond: 3 shared pairs (6 electrons). Written as a triple line: N≡N

Example 3.3 — Drawing a Covalent Bond

How do two hydrogen atoms form H_2?

Each hydrogen atom has 1 valence electron. To fill its first shell (capacity: 2), each needs 1 more electron.

H• + •H → H:H → written as H–H

By sharing the pair, each hydrogen "sees" 2 electrons in its first shell — satisfied.

Example 3.4 — Water (H_2O)

Oxygen has 6 valence electrons. How does it bond with hydrogen to form water?

Oxygen needs 2 more electrons to reach an octet. Two hydrogen atoms each need 1 more. Oxygen forms one single bond with each hydrogen:

H–O–H

In water, oxygen has 2 bonding pairs and 2 lone pairs (non-bonding pairs of electrons). The electron geometry is tetrahedral, but the molecular shape is **bent** (V-shaped). This bent shape, combined with oxygen's electron-pulling power, makes water a **polar molecule** — a property responsible for most of water's remarkable chemical behavior.

3.4 Polarity and Electronegativity

Electronegativity is the measure of how strongly an atom attracts shared electrons toward itself. It increases going up and to the right on the periodic table. Fluorine (F) is the most electronegative element.

When two atoms of different electronegativity share electrons, the electrons spend more time near the more electronegative atom. This creates a **polar covalent bond** — one end has a slight negative charge ($\delta-$) and the other a slight positive charge ($\delta+$).

Nonpolar covalent bond: Equal or nearly equal electronegativity (e.g., H_2, O_2, N_2). Electrons shared symmetrically.

Polar covalent bond: Significant difference in electronegativity. Electrons pulled toward the more electronegative atom.

Ionic bond: Very large electronegativity difference (typically > 1.7). Electrons effectively transferred completely.

Example 3.5 — Classifying Bonds by Polarity

Classify the bonds in: (a) Cl_2; (b) HCl; (c) NaCl.

(a) Cl_2: Two identical atoms → zero electronegativity difference → **nonpolar covalent**

(b) HCl: Cl is more electronegative than H → electrons pulled toward Cl → **polar covalent**. Cl end is $\delta-$; H end is $\delta+$.

(c) NaCl: Large electronegativity difference → **ionic**. Cl effectively takes Na's electron.

3.5 Metallic Bonding

In metals, the bonding model is different from both ionic and covalent. Metal atoms release their valence electrons into a shared "sea" of electrons that moves freely throughout the material. The positive metal ions sit in this electron sea, held together by the attraction between the cations and the surrounding electrons.

This **metallic bonding** model explains metal properties directly: - **Electrical conductivity:** The free electrons move readily under an applied voltage. - **Thermal conductivity:** Free electrons transfer thermal energy rapidly. - **Malleability/ductility:** The non-directional bonding allows layers of atoms to slide without breaking bonds. - **Metallic luster:** Free electrons interact with light and reflect it.

Physics Bridge 3: Chemical Bonds and Physical Forces

The forces in chemical bonds are fundamentally **electromagnetic** — the same force that physics studies at a larger scale. Ionic bonding is the attraction between point charges (Coulomb's law). Covalent bonding involves quantum mechanical effects that physics describes in the study of the hydrogen atom. Metallic bonding's "electron sea" is the basis of classical electromagnetism and circuit theory.

Polarity — which you just learned — reappears in physics as the concept of an **electric dipole**: a separation of positive and negative charge. Dipoles interact with electric fields, and understanding molecular polarity is essential for understanding how capacitors, dielectrics, and polar molecules in solution behave under applied electric fields.

Chapter 3 Review Questions

1. What is the octet rule and what is the main exception to it?
2. Explain, using the concept of valence electrons, why magnesium typically forms a 2+ ion.
3. Draw the electron-dot structure (Lewis structure) for: (a) F_2; (b) NH_3 (nitrogen bonds to three hydrogens, nitrogen has 5 valence electrons); (c) CO_2 (carbon forms double bonds with each oxygen).
4. Classify the bond in each pair as ionic, polar covalent, or nonpolar covalent: (a) K and Br; (b) C and H; (c) O and O.
5. Explain, using the metallic bonding model, why copper is used for electrical wiring.
6. Water (H_2O) is a polar molecule but carbon dioxide (CO_2) is nonpolar overall. Explain why (hint: consider the shape of each molecule and the direction of the polar bonds).

Chapter Four: Chemical Formulas, Nomenclature, and Moles

4.1 Chemical Formulas

A **chemical formula** uses element symbols and subscripts to show the composition of a substance.

- **Molecular formula:** Shows the actual number of each type of atom in one molecule. Example: $C_6H_{12}O_6$ (glucose — 6 carbon, 12 hydrogen, 6 oxygen atoms per molecule).
- **Empirical formula:** Shows the simplest whole-number ratio of atoms. Glucose's empirical formula is CH_2O.
- **Structural formula:** Shows how atoms are connected, not just how many. Used in organic chemistry.

4.2 Naming Compounds — Nomenclature

Naming Binary Ionic Compounds (Metal + Nonmetal)

1. Name the cation (metal) first, using the element name unchanged.
2. Name the anion (nonmetal) second, replacing the ending with **-ide**.

Examples: - NaCl → sodium **chlor**ide - CaO → calcium **ox**ide - $MgBr_2$ → magnesium **brom**ide - Al_2S_3 → aluminum **sulf**ide

Transition metals often form multiple ions. Use Roman numerals to indicate the charge: - $FeCl_2$ → iron(II) chloride (Fe^{2+}) - $FeCl_3$ → iron(III) chloride (Fe^{3+})

Naming Binary Covalent Compounds (Nonmetal + Nonmetal)

Use Greek prefixes to indicate the number of atoms of each element. Drop the final vowel of the prefix before a vowel.

Number	Prefix
1	mono- (omit for first element)
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-

The second element gets the **-ide** ending.

Examples: - CO → carbon **mono**xide - CO_2 → carbon **di**oxide - NO_2 → nitrogen **di**oxide - N_2O_4 → **di**nitrogen **tetra**oxide - PCl_5 → phosphorus **penta**chloride

Example 4.1 — Naming Practice

Name: (a) K_2O; (b) $CuCl_2$; (c) N_2O; (d) SF_6

(a) K_2O: Potassium oxide (Group 1 → K^+, so no Roman numeral needed)

(b) $CuCl_2$: Two Cl^- ions (total charge −2) → Cu must be +2 → **Copper(II) chloride**

(c) N_2O: Two nitrogen, one oxygen → **Dinitrogen monoxide** (laughing gas)

(d) SF_6: One sulfur, six fluorine → **Sulfur hexafluoride**

4.3 Polyatomic Ions

Some groups of atoms carry a net charge and behave as a unit. These are **polyatomic ions**. The most important ones to memorize:

Ion	Formula	Ion	Formula
Ammonium	NH_4^+	Carbonate	CO_3^{2-}
Hydroxide	OH^-	Bicarbonate	HCO_3^-
Nitrate	NO_3^-	Phosphate	PO_4^{3-}

Ion	Formula	Ion	Formula
Sulfate	SO_4^{2-}	Acetate	$C_2H_3O_2^-$

When a polyatomic ion appears more than once in a formula, it is enclosed in parentheses:

$Ca(NO_3)_2$ = calcium nitrate (one Ca^{2+} ion, two NO_3^- ions) $Al_2(SO_4)_3$ = aluminum sulfate (two Al^{3+} ions, three SO_4^{2-} ions)

4.4 The Mole — Chemistry's Counting Unit

Atoms are unimaginably small. A single grain of sand contains roughly 2×10^{19} silicon atoms. Chemists need a practical unit for counting atoms, molecules, and ions. That unit is the **mole** (mol).

1 mole = 6.022×10^{23} particles (atoms, molecules, ions, or other specified particles)

This number — **Avogadro's number (N_a)** — was chosen so that one mole of any element has a mass in grams equal to its atomic mass in amu. This mass is called the **molar mass** (M) and has units of g/mol.

Examples: - 1 mol of carbon (C) = 12.01 g of C - 1 mol of water (H_2O) = 18.02 g of H_2O ($2 \times 1.008 + 16.00 = 18.02$) - 1 mol of glucose ($C_6H_{12}O_6$) = 180.18 g

Example 4.2 — Mole Calculations

How many grams are in 2.50 mol of NaCl?

Step 1: Find the molar mass of NaCl.

Na: 22.99 g/mol + Cl: 35.45 g/mol = **58.44 g/mol**

Step 2: Convert moles to grams using dimensional analysis.

2.50 mol × (58.44 g / 1 mol) = **146 g**

Example 4.3 — Grams to Moles

How many moles are in 45.0 g of water (H_2O)?

Molar mass of H_2O = 2(1.008) + 16.00 = **18.02 g/mol**

45.0 g × (1 mol / 18.02 g) = **2.50 mol**

Example 4.4 — Moles to Atoms

How many molecules are in 2.50 mol of water?

2.50 mol × (6.022×10^{23} molecules / 1 mol) = **1.51×10^{24} molecules**

Physics Bridge 4: The Mole and Statistical Mechanics

The mole concept connects directly to **statistical mechanics** — the branch of physics that explains the large-scale behavior of gases, liquids, and solids in terms of the average behavior of their enormous numbers of particles.

In physics, you will encounter the **ideal gas law**: PV = nRT, where n is the number of moles. The "n" in that equation is the mole concept you just mastered. The gas constant R (8.314 J/mol·K) connects the mole to energy and temperature.

Avogadro's number also appears in physics as the link between microscopic (per atom) and macroscopic (per mole) quantities. For example, the **Boltzmann constant** ($k_B = 1.381 \times 10^{-23}$ J/K) is related to the ideal gas constant R by: $R = N_a \times k_B$. Every equation in thermodynamics and statistical mechanics that uses the Boltzmann constant (microscopic) has a molar-scale counterpart that uses R — and the bridge between them is always Avogadro's number.

Chapter 4 Review Questions

1. What is the difference between a molecular formula and an empirical formula? Give an example of each.

2. Name the following compounds: (a) Li_2S; (b) SnO_2; (c) N_2O_5; (d) $Ba(OH)_2$.

3. Write formulas for: (a) aluminum nitrate; (b) diphosphorus pentoxide; (c) iron(III) oxide; (d) ammonium sulfate.

4. Calculate the molar mass of: (a) Fe_2O_3; (b) $Ca(NO_3)_2$; (c) $C_{12}H_{22}O_{11}$ (table sugar).

5. A sample contains 0.750 mol of CO_2. What is its mass in grams?

6. How many atoms of nitrogen are present in 5.00 g of NH_3? (Hint: first find moles of NH_3, then moles of N atoms, then use Avogadro's number.)

Chapter Five: Chemical Reactions and Balancing Equations

5.1 Writing Chemical Equations

A **chemical equation** is the shorthand representation of a chemical reaction. It shows: - **Reactants** (starting materials) on the left side - **Products** (substances formed) on the right side - An arrow (→) indicating the direction of the reaction

Example: Hydrogen gas reacts with oxygen gas to form water:

$H_2 + O_2 \rightarrow H_2O$ (unbalanced)

State symbols are often included in parentheses after each formula: - (s) = solid - (l) = liquid - (g) = gas - (aq) = aqueous (dissolved in water)

5.2 Balancing Equations — Conservation of Mass

Dalton's principle that atoms are neither created nor destroyed in chemical reactions means that a balanced equation must have the same number of each type of atom on both sides. We balance equations by adjusting **coefficients** (the numbers in front of formulas) — never by changing subscripts.

Example 5.1 — Balancing a Simple Equation

Balance: $H_2 + O_2 \rightarrow H_2O$

Step 1: Count atoms on each side. Left: 2 H, 2 O | Right: 2 H, 1 O — oxygen is unbalanced.

Step 2: Place a coefficient 2 in front of H_2O to balance oxygen: $H_2 + O_2 \rightarrow 2\ H_2O$ Left: 2 H, 2 O | Right: 4 H, 2 O — now hydrogen is unbalanced.

Step 3: Place a coefficient 2 in front of H_2: $2\ H_2 + O_2 \rightarrow 2\ H_2O$ Left: 4 H, 2 O | Right: 4 H, 2 O ✓ **Balanced.**

Example 5.2 — Balancing a More Complex Equation

Balance the combustion of propane: $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$

Step 1: Start with the most complex molecule. C_3H_8 has 3 C and 8 H.

Step 2: Balance carbon first. Need 3 CO_2: $C_3H_8 + O_2 \rightarrow 3\ CO_2 + H_2O$

Step 3: Balance hydrogen. 8 H requires 4 H_2O: $C_3H_8 + O_2 \rightarrow 3\ CO_2 + 4\ H_2O$

Step 4: Balance oxygen. Right side: 3×2 + 4×1 = 6 + 4 = 10 oxygen atoms → need 5 O_2: $C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$

Check: C: 3=3 ✓, H: 8=8 ✓, O: 10=10 ✓ **Balanced.**

5.3 Types of Chemical Reactions

Chemists recognize five fundamental reaction patterns:

1. Synthesis (Combination)

Two or more reactants combine to form a single product.

Pattern: A + B → AB

Example: $2\ Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$

Sodium and chlorine combine to form sodium chloride.

2. Decomposition

A single reactant breaks down into two or more products.

Pattern: AB → A + B

Example: $2\ H_2O_2(l) \rightarrow 2\ H_2O(l) + O_2(g)$

Hydrogen peroxide decomposes into water and oxygen. This is why hydrogen peroxide bubbles when applied to a cut — the enzyme catalase in blood accelerates the decomposition.

3. Single Replacement (Single Displacement)

One element replaces another element in a compound.

Pattern: A + BC → AC + B

Example: $Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$

Iron displaces copper from copper sulfate solution. The solution turns from blue to pale green, and copper metal deposits on the iron. This works because iron is more reactive than copper.

4. Double Replacement (Double Displacement)

The cations of two ionic compounds switch partners.

Pattern: AB + CD → AD + CB

Example: $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

Silver nitrate and sodium chloride solutions react to form a white precipitate of silver chloride (AgCl) and sodium nitrate remains in solution.

5. Combustion

A fuel reacts with oxygen to produce carbon dioxide and water (if the fuel contains carbon and hydrogen). Combustion releases energy as heat and light.

Example: $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g) + \text{energy}$

Methane (natural gas) burns in oxygen.

Example 5.3 — Classifying Reactions

Classify: (a) $2\ SO_2 + O_2 \rightarrow 2\ SO_3$; (b) $CaCO_3 \rightarrow CaO + CO_2$; (c) $Zn + 2\ HCl \rightarrow ZnCl_2 + H_2$; (d) $2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$

(a) Two reactants → one product: **Synthesis**

(b) One reactant → two products: **Decomposition**

(c) Zinc replaces hydrogen in hydrochloric acid: **Single replacement**

(d) Fuel (butane) + oxygen → CO_2 + H_2O: **Combustion**

5.4 Stoichiometry — Using Balanced Equations Quantitatively

The coefficients in a balanced equation represent **mole ratios** of reactants and products. These ratios are conversion factors used to calculate how much product forms from a given amount of reactant — the art of **stoichiometry**.

Example 5.4 — Mole-to-Mole Stoichiometry

In the reaction $N_2 + 3\ H_2 \rightarrow 2\ NH_3$, how many moles of NH_3 are produced from 4.50 mol of H_2?

The balanced equation gives the ratio: 3 mol H_2 : 2 mol NH_3

4.50 mol H_2 × (2 mol NH_3 / 3 mol H_2) = **3.00 mol NH_3**

Example 5.5 — Mass-to-Mass Stoichiometry

How many grams of water form when 6.00 g of hydrogen gas burns completely?

Balanced equation: $2\ H_2 + O_2 \rightarrow 2\ H_2O$

Step 1: Convert grams of H_2 to moles. Molar mass of H_2 = 2.016 g/mol 6.00 g ÷ 2.016 g/mol = 2.976 mol H_2

Step 2: Use mole ratio. 2 mol H_2 : 2 mol H_2O (ratio = 1:1) 2.976 mol H_2 × (2 mol H_2O / 2 mol H_2) = 2.976 mol H_2O

Step 3: Convert moles of H_2O to grams. Molar mass of H_2O = 18.02 g/mol 2.976 mol × 18.02 g/mol = **53.6 g H_2O**

Physics Bridge 5: Chemical Reactions and Energy Conservation

The **Law of Conservation of Mass** (total mass of reactants = total mass of products) is the chemical expression of a universal principle. In physics, this becomes the **conservation of energy and mass** — in fact, Einstein showed that mass and energy are interconvertible ($E = mc^2$). In ordinary chemical reactions, the energy released or absorbed is so small relative to mc^2 that mass appears to be perfectly conserved.

Combustion reactions are the most direct link between chemistry and the **thermodynamics** you will study in physics. When gasoline burns in an engine, chemical potential energy stored in molecular bonds converts to thermal energy, which pushes pistons, which does mechanical work. Physics describes this energy chain with the laws of thermodynamics; chemistry describes the molecular reactions that release the energy.

Stoichiometry in chemistry corresponds to **conservation laws** in physics: both demand exact accounting of conserved quantities (atoms in chemistry; energy, momentum, and charge in physics).

Chapter 5 Review Questions

1. Balance the following equations:
    a. $Al + O_2 \rightarrow Al_2O_3$
    b. $HCl + Ca(OH)_2 \rightarrow CaCl_2 + H_2O$
    c. $C_2H_6 + O_2 \rightarrow CO_2 + H_2O$
2. Classify each balanced equation in question 1 by reaction type.
3. In the reaction $4\ Fe + 3\ O_2 \rightarrow 2\ Fe_2O_3$ (the rusting of iron), how many moles of Fe_2O_3 form when 2.40 mol of Fe reacts completely with sufficient oxygen?
4. How many grams of CO_2 are produced when 28.0 g of CO burns? ($2\ CO + O_2 \rightarrow 2\ CO_2$)
5. A student claims that by placing a 2 subscript on the product formula to balance an equation, such as writing $H_2 + O_2 \rightarrow H_2O_2$. What is wrong with this approach?

Chapter Six: States of Matter and Gas Laws

6.1 Kinetic Molecular Theory

The **kinetic molecular theory (KMT)** is the model chemists and physicists use to explain the behavior of gases. Its key postulates:

1. Gases consist of tiny particles (atoms or molecules) that are very far apart — the volume occupied by the particles themselves is negligible compared to the total volume of the gas.
2. Gas particles move in constant, random, straight-line motion.
3. Collisions between gas particles and between particles and container walls are **perfectly elastic** — no kinetic energy is lost.
4. There are no significant attractive or repulsive forces between gas particles (ideal behavior).
5. The **average kinetic energy** of gas particles is directly proportional to the absolute temperature (in kelvin).

This fifth postulate is profound: **temperature is a measure of the average kinetic energy of the particles in a substance**.

6.2 Pressure

Pressure is force per unit area. Gas pressure results from countless particle collisions with the walls of a container.

SI unit: Pascal (Pa) = $1\ \text{N/m}^2$

Other common pressure units:

Unit	Equivalent to 1 atm
atmosphere (atm)	1 atm
pascal (Pa)	101,325 Pa
kilopascal (kPa)	101.325 kPa
millimeters of mercury (mmHg)	760 mmHg
torr	760 torr

6.3 The Four Gas Laws

The behavior of gases can be described by four key relationships among pressure (P), volume (V), temperature (T, in kelvin), and amount (n, in moles).

Boyle's Law — P and V (constant T, n)

At constant temperature and amount, pressure and volume are **inversely proportional**:

$P_1V_1 = P_2V_2$

As you squeeze a gas into a smaller volume, the pressure increases proportionally.

Example 6.1 — Boyle's Law

A gas occupies 4.00 L at 2.00 atm. What volume does it occupy at 5.00 atm (temperature constant)?

$P_1V_1 = P_2V_2$ (2.00 atm)(4.00 L) = (5.00 atm)(V_2) V_2 = 8.00 / 5.00 = **1.60 L**

The gas was compressed to a smaller volume, as expected.

Charles's Law — V and T (constant P, n)

At constant pressure and amount, volume and temperature are **directly proportional** (temperature in kelvin):

$V_1/T_1 = V_2/T_2$

Example 6.2 — Charles's Law

A balloon occupies 3.00 L at 27°C. What is its volume at 127°C (pressure constant)?

Convert to kelvin: T_1 = 27 + 273 = 300 K; T_2 = 127 + 273 = 400 K

$V_1/T_1 = V_2/T_2$ 3.00/300 = V_2/400 V_2 = 3.00 × 400/300 = **4.00 L**

Critical habit: Always convert Celsius to kelvin before using any gas law: **K = °C + 273.15**

Gay-Lussac's Law — P and T (constant V, n)

At constant volume and amount, pressure and temperature are **directly proportional**:

$P_1/T_1 = P_2/T_2$

This is why pressure cookers reach higher temperatures, and why a sealed aerosol can in a fire is dangerous — rising temperature increases pressure until the can ruptures.

The Ideal Gas Law — Combining All Four

Combining Boyle's, Charles's, and Gay-Lussac's laws with Avogadro's principle (equal volumes of gases at the same T and P contain the same number of moles) gives the **ideal gas law**:

PV = nRT

Where: - P = pressure in atm (or Pa) - V = volume in liters (or m^3) - n = moles - T = temperature in kelvin - R = **ideal gas constant** = 0.08206 L·atm/mol·K (or 8.314 J/mol·K)

Example 6.3 — Ideal Gas Law

How many moles of nitrogen gas are in a 10.0 L tank at 25°C and 3.50 atm?

T = 25 + 273 = 298 K

PV = nRT → n = PV / RT

n = (3.50 atm × 10.0 L) / (0.08206 L·atm/mol·K × 298 K)

n = 35.0 / 24.45 = **1.43 mol N_2**

6.4 Dalton's Law of Partial Pressures

In a mixture of gases, each gas exerts its own **partial pressure** independently of the others. The total pressure is the sum of all partial pressures:

$P_{total} = P_1 + P_2 + P_3 + \ldots$

Example 6.4 — Partial Pressures

A container holds 0.500 mol N_2, 0.300 mol O_2, and 0.200 mol CO_2 at 300 K in a 10.0 L vessel. What is the total pressure?

Total moles: n = 0.500 + 0.300 + 0.200 = 1.000 mol

$P_{total} = nRT/V = (1.000)(0.08206)(300) / 10.0 =$ **2.46 atm**

(Alternatively, find each partial pressure and add: $P(N_2)$ = 1.23 atm, $P(O_2)$ = 0.738 atm, $P(CO_2)$ = 0.492 atm, sum = 2.46 atm.)

Physics Bridge 6: Gas Laws Are Physics Laws

The ideal gas law PV = nRT is simultaneously a chemistry equation and a fundamental physics equation. In physics, it appears in exactly this form in **thermodynamics** and **statistical mechanics**.

Temperature and kinetic energy: Physics defines temperature through the equipartition theorem: the average kinetic energy per molecule = (3/2)k_BT, where k_B is the Boltzmann constant. The absolute zero of temperature (0 K = −273.15°C) is the temperature at which molecular motion would cease — a concept you already touched on through Charles's Law (the volume of an ideal gas extrapolates to zero at 0 K).

Pressure and momentum: Physics explains gas pressure through Newton's second law: when a gas molecule bounces off a wall, it transfers momentum to the wall, and the rate of momentum transfer is force. Summing over all molecules gives the pressure formula, which leads directly to Boyle's Law from first principles.

The ideal gas law in engineering: Every heat engine (steam engines, internal combustion engines, jet engines, rocket engines) operates by expanding a hot gas to do work. The ideal gas law is the starting equation for understanding all of these devices. In physics, you will use PV = nRT to derive the efficiency of ideal heat engines (Carnot cycles).

Chapter 6 Review Questions

1. State the five postulates of the kinetic molecular theory.

2. Convert 750.0 mmHg to atm and to kPa.

3. A gas occupies 2.50 L at 1.20 atm. What pressure is needed to compress it to 1.00 L at constant temperature?

4. A gas sample occupies 5.00 L at 0°C. What volume does it occupy at 100°C at the same pressure?

5. A 2.00 L container holds 0.100 mol of a gas at 300 K. What is the pressure in atm?

6. Explain, using the kinetic molecular theory, why the pressure of a gas in a sealed container increases when the container is heated.

Chapter Seven: Solutions, Concentration, and Acids and Bases

7.1 Solutions

A **solution** is a homogeneous mixture in which one substance (the **solute**) is dissolved in another (the **solvent**). The solvent is the component present in the larger amount; water is by far the most common solvent.

"Like dissolves like" is the guiding rule of solubility: polar solvents dissolve polar solutes, and nonpolar solvents dissolve nonpolar solutes.

- Water (polar) dissolves salt (ionic) and sugar (polar molecular compound). ✓
- Water does not dissolve oil (nonpolar). ✗
- Gasoline (nonpolar) dissolves motor oil (nonpolar). ✓

Electrolytes and Nonelectrolytes

- **Electrolyte:** Dissolves to produce ions in solution — conducts electricity. Strong electrolytes (e.g., NaCl, HCl) dissociate completely. Weak electrolytes (e.g., acetic acid) partially dissociate.
- **Nonelectrolyte:** Dissolves without producing ions — does not conduct electricity. Sugar in water is a nonelectrolyte.

7.2 Concentration — Molarity

Molarity (M) is the most common concentration unit in chemistry:

Molarity (M) = moles of solute / liters of solution

Example 7.1 — Calculating Molarity

What is the molarity of a solution prepared by dissolving 5.85 g of NaCl in enough water to make 250.0 mL of solution?

Step 1: Convert grams to moles. Molar mass NaCl = 58.44 g/mol 5.85 g ÷ 58.44 g/mol = 0.1001 mol NaCl

Step 2: Convert mL to L. 250.0 mL = 0.2500 L

Step 3: Calculate molarity. M = 0.1001 mol / 0.2500 L = **0.400 M**

Example 7.2 — Dilution

A **dilution** occurs when solvent is added to a solution. The amount of solute stays constant:

$M_1V_1 = M_2V_2$

What volume of 12.0 M HCl is needed to prepare 500.0 mL of 0.500 M HCl?

$M_1V_1 = M_2V_2$ $(12.0\ M)(V_1) = (0.500\ M)(500.0\ mL)$ $V_1 = 250 / 12.0 =$ **20.8 mL**

Measure 20.8 mL of concentrated HCl and add it to enough water to reach a total volume of 500.0 mL.

7.3 Acids and Bases — Arrhenius Definition

Arrhenius acid: A substance that produces H^+ ions (protons) when dissolved in water. - $HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$

Arrhenius base: A substance that produces OH^- ions when dissolved in water. - $NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$

7.4 The Brønsted-Lowry Definition

More general: an **acid** is a proton (H^+) *donor*; a **base** is a proton *acceptor*.

In the reaction between HCl and water: $HCl + H_2O \rightarrow H_3O^+ + Cl^-$

HCl donates H^+ to water → HCl is the **acid** H_2O accepts H^+ → H_2O acts as the **base**

This definition explains why water can act as both an acid and a base (it is **amphoteric**).

7.5 Strong and Weak Acids and Bases

Strong acids dissociate *completely* in water. The seven common strong acids: HCl, HBr, HI, HNO_3, H_2SO_4, $HClO_4$, $HClO_3$.

Weak acids dissociate *partially*. Acetic acid (vinegar) is a weak acid: $CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$ (only about 1% dissociates at typical concentrations)

Strong bases: NaOH, KOH, $Ca(OH)_2$, and other soluble metal hydroxides.

Weak bases: NH_3 and most organic amines.

7.6 The pH Scale

The concentration of H^+ ions in a solution determines its acidity. Because this concentration spans many orders of magnitude, chemists use the logarithmic **pH scale**:

$pH = -\log[H^+]$

pH	Character	Example
0	Strongly acidic	Battery acid
1	Strongly acidic	Stomach acid
3	Acidic	Lemon juice
7	Neutral	Pure water
9	Basic	Baking soda
12	Strongly basic	Bleach
14	Strongly basic	Drain cleaner

Example 7.3 — pH Calculations

(a) What is the pH of a solution with $[H^+] = 0.0100$ M? (b) What is $[H^+]$ if pH = 4.50?

(a) pH = $-\log(0.0100) = -\log(10^{-2}) =$ **2.00**

(b) $[H^+] = 10^{(-pH)} = 10^{(-4.50)} =$ **3.16×10^{-5} M**

7.7 Neutralization

When an acid and a base react, they produce a **salt** and water. This is an **acid-base neutralization**:

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

At the point of neutralization (equal moles of H^+ and OH^- have reacted), the pH is 7 for strong acid + strong base reactions.

Physics Bridge 7: Electricity and Ion Concentration

The concept of **electrolytes** — solutions that conduct electricity because they contain mobile ions — is the chemical foundation of **electrochemistry** in physics.

Ionic concentration determines electrical conductivity. In physics, you will study how electric current flows through liquids (electrolysis), which requires understanding that current is carried by ions in solution, not by free electrons (as in metals).

pH and logarithms provide practice with a mathematical tool used throughout physics: logarithmic scales. The **decibel scale** for sound intensity, the **Richter scale** for earthquake magnitude, and the **stellar magnitude scale** in astronomy are all logarithmic, exactly like pH.

The **electric potential** (voltage) in electrochemical cells — the chemistry of batteries — is related to the concentration of ions by the Nernst equation, which uses the same logarithmic relationship as pH. Every battery in every device you own operates on principles that combine acid-base chemistry with electromagnetic theory.

Chapter 7 Review Questions

1. Explain, using the principle of "like dissolves like," why iodine (I_2) dissolves readily in hexane (a nonpolar liquid) but not in water.

2. What is the molarity of a solution prepared by dissolving 25.0 g of glucose ($C_6H_{12}O_6$, M = 180.18 g/mol) in 500.0 mL of solution?

3. What volume of 6.00 M NaOH is needed to prepare 250 mL of 0.150 M NaOH?

4. Classify each as a strong or weak acid/base: (a) HNO_3; (b) CH_3COOH; (c) KOH; (d) NH_3.

5. Calculate the pH of: (a) $[H^+] = 0.0500$ M; (b) $[H^+] = 2.5 \times 10^{-9}$ M. Is each acidic, neutral, or basic?

6. Write the balanced neutralization equation between sulfuric acid (H_2SO_4) and potassium hydroxide (KOH), including state symbols.

Chapter Eight: Thermochemistry — Heat and Chemical Energy

8.1 Energy, Heat, and Work

Energy is the capacity to do work or transfer heat. Chemistry deals primarily with two forms of energy:

Potential energy (chemical): Energy stored in chemical bonds. When bonds break and new ones form, energy is released or absorbed.

Thermal energy (heat): Energy associated with the random motion of particles. The more energetic the motion, the higher the temperature.

Heat (q): Energy that flows between objects because of a temperature difference. Heat always flows from higher temperature to lower temperature (this is the Second Law of Thermodynamics, previewed here and central to physics).

Work (w): In chemistry, work done by or on a gas when it expands or is compressed against external pressure.

First Law of Thermodynamics: Energy is conserved. The change in internal energy of a system equals heat added to the system minus work done by the system:

$$\Delta E = q - w$$

8.2 Exothermic and Endothermic Reactions

Exothermic reaction: Releases energy (usually as heat) to the surroundings. The products have less chemical potential energy than the reactants. The surroundings warm up. ΔH is negative.

Examples: combustion, neutralization, rusting, many synthesis reactions.

Endothermic reaction: Absorbs energy from the surroundings. The products have more chemical potential energy than the reactants. The surroundings cool down. ΔH is positive.

Examples: photosynthesis, dissolving ammonium nitrate in water (cold packs), decomposition reactions.

8.3 Enthalpy — ΔH

Enthalpy (H) is a measure of the total heat content of a system at constant pressure. The **change in enthalpy (ΔH)** is the heat released or absorbed during a reaction at constant pressure (which is true for most laboratory reactions open to the atmosphere):

ΔH = H_products – H_reactants

- If ΔH < 0: reaction is **exothermic** (releases heat)
- If ΔH > 0: reaction is **endothermic** (absorbs heat)

Thermochemical equation: A balanced equation that includes ΔH.

$CH_4(g) + 2\,O_2(g) \rightarrow CO_2(g) + 2\,H_2O(l)\ \Delta H = -890\ kJ$

This means burning 1 mole of methane releases 890 kJ of heat.

8.4 Specific Heat Capacity

Different substances require different amounts of heat to change temperature. **Specific heat capacity (c)** is the amount of heat required to raise 1 gram of a substance by 1°C (or 1 K).

Substance	Specific Heat (J/g·°C)
Water (l)	4.184
Ice (s)	2.09
Aluminum	0.900
Iron	0.449
Copper	0.385

Water's high specific heat capacity explains why coastal cities have milder climates than inland cities, and why water is used in engine cooling systems.

The heat equation:

q = m × c × ΔT

Where: - q = heat (joules) - m = mass (grams) - c = specific heat capacity (J/g·°C) - ΔT = change in temperature (°C or K)

Example 8.1 — Specific Heat

How much heat is required to raise the temperature of 250.0 g of water from 25.0°C to 100.0°C?

$\Delta T = 100.0 - 25.0 = 75.0°C$

$q = m \times c \times \Delta T = 250.0\ g \times 4.184\ J/g·°C \times 75.0°C$

$q =$ **78,450 J = 78.5 kJ**

Example 8.2 — Calorimetry

A 50.0 g piece of copper at 90.0°C is dropped into 100.0 g of water at 20.0°C. What is the final temperature? (Assume no heat is lost to surroundings.)

Heat lost by copper = heat gained by water $q_{Cu} = -q_{water}$ (the sign convention: heat leaving copper = heat entering water)

$m_{Cu} \times c_{Cu} \times \Delta T_{Cu} = -(m_{water} \times c_{water} \times \Delta T_{water})$

$50.0 \times 0.385 \times (T_f - 90.0) = -(100.0 \times 4.184 \times (T_f - 20.0))$

$19.25(T_f - 90.0) = -418.4(T_f - 20.0)$

$19.25\ T_f - 1732.5 = -418.4\ T_f + 8368$

$437.65\ T_f = 10100.5$

$T_f =$ **23.1°C**

The copper cooled dramatically; the water warmed only slightly — because water has a much higher heat capacity.

8.5 Hess's Law

Hess's Law: The total enthalpy change for a reaction is the same regardless of whether the reaction occurs in one step or many steps. Enthalpy is a **state function** — it depends only on the initial and final states, not the path.

This allows us to calculate ΔH for reactions that are difficult to measure directly by combining reactions whose ΔH values are known.

Example 8.3 — Hess's Law

Calculate ΔH for: $C(s) + \frac{1}{2}\ O_2(g) \rightarrow CO(g)$

Given: (1) $C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H_1 = -393.5$ kJ (2) $CO(g) + \frac{1}{2}\ O_2(g) \rightarrow CO_2(g)$ $\Delta H_2 = -283.0$ kJ

To get the target equation, keep equation (1) as is, and reverse equation (2):

(1) $C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H = -393.5$ kJ (2 reversed) $CO_2(g) \rightarrow CO(g) + \frac{1}{2}\ O_2(g)$ $\Delta H = +283.0$ kJ

Add the two equations (CO_2 cancels): $C(s) + O_2(g) - \frac{1}{2}\ O_2(g) \rightarrow CO(g)$ $C(s) + \frac{1}{2}\ O_2(g) \rightarrow CO(g)$

$\Delta H_{total} = -393.5 + 283.0 =$ **−110.5 kJ**

Physics Bridge 8: Thermodynamics Unified

Chapter 8 is where chemistry and physics most directly overlap. **Thermodynamics** is a shared subject.

The **First Law of Thermodynamics (ΔE = q − w)** introduced here is the same law that appears throughout physics. In physics, you will apply it to heat engines (where work is the piston moving), refrigerators, and electrical systems.

Specific heat capacity reappears in physics as one of the key material properties in **calorimetry** and **thermal equilibrium**. The same equation q = mcΔT is used in physics to calculate heat flow in everything from cooling electronics to designing building insulation.

State functions in chemistry (enthalpy, energy) correspond to **path-independent quantities** in physics. In physics, you will learn that the work done by a conservative force (gravity, springs) is also path-independent — Hess's Law and gravitational potential energy both reflect this deeper mathematical structure.

Entropy — introduced briefly as a physics concept — explains *why* reactions proceed: systems move toward states of lower energy and higher disorder. This becomes the **Second Law of Thermodynamics** in physics.

Chapter 8 Review Questions

1. A hand warmer undergoes a chemical reaction that heats up. Is this exothermic or endothermic? Is ΔH positive or negative?

2. How much heat is released when 500.0 g of iron cools from 200.0°C to 25.0°C? (c_Fe = 0.449 J/g·°C)

3. A reaction releases 245 kJ per mole of reactant. If 5.00 g of the reactant (molar mass = 50.0 g/mol) reacts, how much heat is released?

4. Using Hess's Law, calculate ΔH for the reaction $S(s) + O_2(g) \rightarrow SO_2(g)$, given:

   - $S(s) + 3/2\ O_2(g) \rightarrow SO_3(g)$ ΔH = −396 kJ
   - $SO_2(g) + ½\ O_2(g) \rightarrow SO_3(g)$ ΔH = −98 kJ

5. Explain, using specific heat capacity values, why metals feel colder to the touch than wood at room temperature, even though both are at the same temperature.

Chapter Nine: Reaction Rates and Chemical Equilibrium

9.1 Reaction Rates

Reaction rate is how quickly reactants are consumed or products are formed, typically expressed as the change in concentration per unit time:

Rate = Δ[concentration] / Δtime

Units: mol/L·s (or $mol \cdot L^{-1} \cdot s^{-1}$)

Factors Affecting Reaction Rate

1. Concentration: Higher concentration → more frequent collisions → faster rate.

2. Temperature: Higher temperature → more kinetic energy → more molecules exceed the activation energy → faster rate. Roughly, a 10°C increase doubles the reaction rate.

3. Surface area: For reactions involving solids, more surface area → more collision sites → faster rate. Powdered iron rusts much faster than a solid iron bar.

4. Catalyst: A substance that increases the reaction rate *without being consumed*. Catalysts provide an alternative pathway with lower activation energy.

5. Nature of reactants: Reactions between ions in solution (ionic reactions) occur almost instantaneously. Reactions that require breaking strong covalent bonds are slower.

9.2 Collision Theory and Activation Energy

For a reaction to occur, molecules must **collide** with: 1. Sufficient energy to overcome the **activation energy (E_a)** — the minimum energy required to break existing bonds and begin forming new ones. 2. The correct **orientation** — molecules must collide in a geometry that allows the reaction to proceed.

Energy diagrams illustrate this:

- **Exothermic reaction:** Products are at a lower energy than reactants. The difference is ΔH (negative). A hill (the activation energy barrier) must be crossed.
- **Endothermic reaction:** Products are at a higher energy. ΔH is positive. The activation energy is even larger.

A **catalyst** lowers the activation energy hill — it opens a shorter path between the same reactants and products.

9.3 Chemical Equilibrium

Many reactions are **reversible**: products can react to re-form reactants. A double arrow (⇌) indicates reversibility:

$N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)$

Initially, only the forward reaction occurs. As products build up, the reverse reaction begins. Eventually, the forward and reverse rates become equal — a state of **dynamic equilibrium**. At equilibrium: - Concentrations of reactants and products remain *constant* (but not necessarily equal). - Both the forward and reverse reactions continue at equal rates. - The system is not static — it is dynamically balanced.

9.4 The Equilibrium Constant (K)

For the general reaction: $aA + bB \rightleftharpoons cC + dD$

The equilibrium constant expression is:

K = [C]^c [D]^d / [A]^a [B]^b

(Square brackets denote molar concentration; pure solids and liquids are excluded.)

- **K >> 1:** Equilibrium lies far to the right — products predominate.
- **K << 1:** Equilibrium lies far to the left — reactants predominate.
- **K ≈ 1:** Significant amounts of both reactants and products are present.

Example 9.1 — Writing K Expressions

Write the equilibrium expression for: $2\,SO_2(g) + O_2(g) \rightleftharpoons 2\,SO_3(g)$

$K = [SO_3]^2 / ([SO_2]^2\,[O_2])$

Example 9.2 — Calculating K

At 500 K, an equilibrium mixture contains: $[SO_2] = 0.200$ M, $[O_2] = 0.100$ M, $[SO_3] = 0.400$ M. Calculate K.

$K = (0.400)^2 / ((0.200)^2 \times 0.100)$

$K = 0.160 / (0.0400 \times 0.100)$

$K = 0.160 / 0.00400 =$ **40.0**

K > 1, so at equilibrium products are favored.

9.5 Le Chatelier's Principle

Le Chatelier's Principle: If a system at equilibrium is disturbed (by changing concentration, temperature, or pressure), the equilibrium shifts in the direction that *partially counteracts* the disturbance.

Concentration Changes

- Add more reactant → equilibrium shifts *right* (toward products) to consume the added reactant.
- Remove product → equilibrium shifts *right* to replenish the product.
- Add more product → equilibrium shifts *left*.

Temperature Changes

- Treat heat as a reactant (endothermic) or product (exothermic).
- Increase temperature of exothermic reaction → shifts *left* (removes the added "heat" by reversing the reaction).
- Decrease temperature → shifts *right* (generates heat).

Pressure Changes (Gas-Phase Reactions)

- Increase pressure → shifts toward the side with fewer moles of gas.
- Decrease pressure → shifts toward the side with more moles of gas.

Example 9.3 — Applying Le Chatelier's Principle

For the reaction: $N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)$ $\Delta H = -92$ kJ

Predict the effect of: (a) adding more H_2; (b) increasing temperature; (c) increasing pressure.

(a) Adding H_2 (a reactant) → system shifts *right* → more NH_3 produced.

(b) Increasing temperature adds energy. The reaction is exothermic ($\Delta H < 0$), so heat is a product. Adding heat → shifts *left* → more N_2 and H_2, less NH_3. (This is why high temperature hurts ammonia yield — the Haber process for making fertilizer must balance rate vs. yield.)

(c) Left side: 1 + 3 = 4 mol gas. Right side: 2 mol gas. Increasing pressure → shifts *right* (toward fewer gas moles) → more NH_3.

Physics Bridge 9: Dynamics, Equilibrium, and Stability

Chemical equilibrium maps directly onto **mechanical equilibrium** in physics. In physics, a system in equilibrium has balanced forces — push it and it resists, returning toward equilibrium. In chemistry, a system in equilibrium has balanced reaction rates — disturb it and it resists, shifting partially back.

Le Chatelier's Principle is a special case of the more general physical principle of **negative feedback**: when a system is perturbed from equilibrium, a response arises that opposes the perturbation. Negative feedback stabilizes systems ranging from chemical reactions to thermostats to control systems in engineering to biological hormonal regulation.

Activation energy in chemistry corresponds to the concept of an **energy barrier** in physics — the energy required to move from one stable configuration to another. In physics, you will see this in barrier tunneling (quantum mechanics), phase transitions (thermodynamics), and transition state theory in physical chemistry.

Rate equations and their mathematical form (concentration changes over time) directly connect to **differential equations** in physics — the language used to describe everything from projectile motion to electrical circuits to wave propagation.

Chapter 9 Review Questions

1. List five factors that affect the rate of a chemical reaction and explain the mechanism by which each factor influences rate.
2. What is activation energy? Sketch an energy diagram for both an exothermic and an endothermic reaction, labeling the reactants, products, activation energy, and ΔH.

3. Write the equilibrium constant expression for: $CO(g) + 3\ H_2(g) \rightleftharpoons CH_4(g) + H_2O(g)$

4. For the reaction $H_2(g) + I_2(g) \rightleftharpoons 2\ HI(g)$, K = 55.3 at 700 K. At equilibrium, $[H_2]$ = 0.100 M and $[I_2]$ = 0.100 M. Calculate [HI].

5. For the reaction: $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$ ΔH = +88 kJ, predict the direction of equilibrium shift when: (a) Cl_2 is added; (b) temperature is decreased; (c) volume is decreased (pressure increased).

Chapter Ten: Electrochemistry, Organic Chemistry, and the Bridge to Physics

10.1 Oxidation and Reduction — Redox Reactions

Many of the most important reactions in chemistry and physics involve the transfer of electrons. These are called **oxidation-reduction (redox) reactions**.

Oxidation: Loss of electrons. **"OIL"** — Oxidation Is Loss. **Reduction:** Gain of electrons. **"RIG"** — Reduction Is Gain.

Memory device: OIL RIG

Oxidation and reduction always occur together — if one substance loses electrons, another must gain them. The substance that is oxidized causes another to be reduced (it is the **reducing agent**). The substance that is reduced causes another to be oxidized (it is the **oxidizing agent**).

Oxidation Numbers

To track electrons in redox reactions, we assign **oxidation numbers** (or oxidation states) — hypothetical charges assigned to atoms.

Key rules: 1. Free elements have oxidation number = 0. 2. In a compound, Group 1 metals = +1; Group 2 = +2. 3. Hydrogen = +1 (except in metal hydrides where H = −1). 4. Oxygen = −2 (except in peroxides where O = −1). 5. The sum of oxidation numbers = charge on the species (0 for neutral compounds).

Example 10.1 — Identifying Redox

Identify what is oxidized and reduced in: $Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$

Fe goes from 0 (free element) to +2 (in $FeSO_4$) → loses 2 electrons → **Fe is oxidized** (Fe is the reducing agent).

Cu goes from +2 (in $CuSO_4$) to 0 (free metal) → gains 2 electrons → **Cu is reduced** (Cu^{2+} is the oxidizing agent).

10.2 Electrochemical Cells — Batteries

An **electrochemical cell** uses a spontaneous redox reaction to produce electrical energy (a **galvanic cell** or **voltaic cell**), or uses electrical energy to drive a non-spontaneous redox reaction (an **electrolytic cell**).

Galvanic (voltaic) cells are the chemistry inside every battery: - **Anode:** The electrode where oxidation occurs. Electrons are released here and flow through the external circuit. - **Cathode:** The electrode where reduction occurs. Electrons arrive here from the external circuit. - A **salt bridge** or porous membrane allows ions to flow between the two half-cells, maintaining electrical neutrality.

Cell voltage (electromotive force, EMF): The "push" driving electrons through the external circuit. Measured in volts (V). Standard cell potentials (E°) can be calculated from standard reduction potentials of the two half-reactions:

E°_cell = E°_cathode − E°_anode

Example 10.2 — Cell Voltage

Calculate the standard cell voltage for a cell with Cu^{2+}/Cu at the cathode and Zn^{2+}/Zn at the anode.

Standard reduction potentials: - $Cu^{2+} + 2e^- \rightarrow Cu$ $E^\circ = +0.34$ V - $Zn^{2+} + 2e^- \rightarrow Zn$ $E^\circ = -0.76$ V

$E^\circ_{cell} = (+0.34) - (-0.76) =$ **+1.10 V**

This is the voltage of a Daniell cell — the basis of early batteries and still used in electrochemistry demonstrations. A positive E° confirms the reaction is spontaneous.

10.3 Electrolysis — Driving Reactions with Electricity

Electrolysis is the reverse of a galvanic cell: electrical energy forces a non-spontaneous redox reaction to occur. Industrial applications include: - **Electroplating:** Depositing a thin metal layer on an object (e.g., chrome plating on car parts) - **Aluminum production:** Al_2O_3 is electrolyzed to produce pure aluminum metal — the only commercially viable method - **Chlor-alkali process:** Electrolysis of saltwater produces Cl_2 gas, H_2 gas, and NaOH — three industrial chemicals from a single process.

10.4 Introduction to Organic Chemistry

Organic chemistry is the chemistry of carbon-containing compounds. Carbon's ability to form four bonds, bond to itself in chains and rings, and bond to hydrogen, oxygen, nitrogen, and many other elements makes it the basis of an almost limitless variety of molecules — including all molecules of life.

Hydrocarbons

The simplest organic compounds contain only carbon and hydrogen:

Alkanes: Single bonds only. General formula: C_nH_{2n+2}. Named with the suffix **-ane**. - CH_4: methane (natural gas) - C_2H_6: ethane - C_3H_8: propane - C_4H_{10}: butane

Alkenes: Contain one or more C=C double bonds. General formula: C_nH_{2n}. Suffix: **-ene**. - C_2H_4: ethene (ethylene — used to make polyethylene plastic)

Alkynes: Contain a C≡C triple bond. General formula: C_nH_{2n-2}. Suffix: **-yne**. - C_2H_2: ethyne (acetylene — used in welding torches)

Aromatic hydrocarbons: Contain benzene rings (C_6H_6 with alternating single and double bonds delocalized around the ring). Benzene is the simplest example.

Functional Groups

A **functional group** is a specific arrangement of atoms that gives organic molecules their characteristic chemical properties.

Functional Group	Structure	Class	Example
Hydroxyl	–OH	Alcohol	Ethanol (C_2H_5OH)
Carboxyl	–COOH	Carboxylic acid	Acetic acid (CH_3COOH)
Amino	–NH_2	Amine / Amino acid	Glycine
Carbonyl (aldehyde)	–CHO	Aldehyde	Formaldehyde
Carbonyl (ketone)	–C(=O)–	Ketone	Acetone
Ester	–COO–	Ester	Fats and oils

Polymers

Polymers are giant molecules made by linking many small repeating units (**monomers**) together.

- **Polyethylene:** Made from ethylene monomers ($CH_2{=}CH_2)_n$ → millions of $CH_2–CH_2$ units linked into long chains. Used for plastic bags and bottles.
- **Nylon:** A polyamide — monomers linked through amide bonds. Strong and flexible.
- **DNA:** A biological polymer — the monomer is a nucleotide. The double helix stores genetic information through the specific sequence of four types of nucleotides.

10.5 Nuclear Chemistry — A Gateway to Physics

Nuclear reactions involve changes to the *nucleus* itself — unlike chemical reactions, which involve only the electron cloud.

Radioactive decay: Unstable nuclei emit radiation spontaneously. - **Alpha (α) decay:** Emits $^{2}{}_{2}He$ (a helium-4 nucleus). Loses 2 protons and 2 neutrons. - **Beta (β) decay:** A neutron converts to a proton, emitting a fast electron (beta particle). - **Gamma (γ) radiation:** High-energy electromagnetic radiation emitted when an excited nucleus relaxes.

Half-life: The time required for half the atoms in a radioactive sample to decay. Carbon-14 has a half-life of 5,730 years — the basis of radiocarbon dating.

Example 10.3 — Half-Life

A sample of carbon-14 initially contains 80.0 g. How much remains after 3 half-lives (3 × 5,730 = 17,190 years)?

After 1 half-life: 80.0 / 2 = 40.0 g After 2 half-lives: 40.0 / 2 = 20.0 g After 3 half-lives: 20.0 / 2 = **10.0 g**

Nuclear fission: A heavy nucleus (e.g., U-235) splits into two smaller nuclei, releasing enormous energy. The basis of nuclear power plants and atomic bombs.

Nuclear fusion: Two light nuclei combine to form a heavier nucleus, releasing even more energy. The process that powers the Sun and all stars.

Einstein's equation connecting mass and energy: > **$E = mc^2$**

where $c = 2.998 \times 10^8$ m/s. Even a tiny mass converts to an enormous amount of energy — a 1 kg mass converted entirely to energy would release 9×10^{16} J, equivalent to about 20 million tons of TNT.

Physics Bridge 10: Chemistry as the Foundation of Physics

This final chapter has introduced three topics that open directly onto the major branches of physics:

Electrochemistry → Electromagnetism: The voltage of an electrochemical cell is measured in volts — the same unit as electric potential in physics. Ohm's Law (V = IR), circuits, and electromagnetic induction all build on the concept of electrical potential established here. The oxidation-reduction electron transfer in a battery is the chemical source of the electrical energy that physics then analyzes in terms of fields, forces, and currents.

Organic chemistry → Materials science and biophysics: The polymers introduced here — from polyethylene to DNA — are subjects of **materials science** (a field combining physics, chemistry, and engineering) and **biophysics** (applying physics to biological systems). The physical properties of materials — electrical conductivity, mechanical strength, thermal expansion — emerge from the molecular structures and bonds that chemistry describes.

Nuclear chemistry → Nuclear physics: The half-life equation, radioactive decay, fission, fusion, and $E = mc^2$ are the first steps into **nuclear physics** — one of the great achievements of twentieth-century science. The binding energy of nuclei, the strong nuclear force that holds protons and neutrons together, and the processes that power stars are among the deepest topics in physics. You have taken the first steps toward them by understanding the difference between chemical and nuclear reactions, and by knowing what protons and neutrons are and where they live.

Chapter 10 Review Questions

1. Define oxidation and reduction. In the reaction $2\ Al + 3\ Cl_2 \rightarrow 2\ AlCl_3$, identify what is oxidized and what is reduced, and identify the oxidizing and reducing agents.

2. A galvanic cell is constructed with Mg^{2+}/Mg ($E° = -2.37$ V) as the anode and Ag^{+}/Ag ($E° = +0.80$ V) as the cathode. (a) Write the half-reactions and overall cell reaction. (b) Calculate $E°_{cell}$. (c) Is the reaction spontaneous?

3. Name the organic compounds: (a) C_3H_8; (b) C_4H_8 (one double bond); (c) C_2H_2.

4. What is a functional group? Identify the functional group in: (a) CH_3OH; (b) CH_3COOH; (c) CH_3NH_2.

5. A radioactive sample has a half-life of 20.0 minutes. If the initial mass is 160.0 g, how much remains after 80.0 minutes?

6. Calculate the energy released if 0.0100 kg of matter is completely converted to energy using $E = mc^2$ ($c = 3.00 \times 10^8$ m/s). Express your answer in joules and compare it to the energy released by burning 1 kg of coal (approximately 3.0×10^7 J).

Comprehensive Review and Looking Ahead

Summary of the Ten Chapters

Chapter	Core Concept	Key Skills
1	Nature of matter; states; physical vs. chemical change; measurement	Scientific notation; significant figures
2	Atomic structure; periodic table; electron configuration	Proton/neutron/electron counts; reading the periodic table
3	Ionic, covalent, and metallic bonding; polarity	Lewis structures; classifying bonds
4	Formulas; nomenclature; the mole	Naming compounds; mole-gram-atom conversions
5	Chemical equations; reaction types; stoichiometry	Balancing equations; mass-to-mass calculations
6	Kinetic molecular theory; gas laws; ideal gas law	Boyle's, Charles's, Gay-Lussac's, ideal gas law calculations
7	Solutions; concentration; acids/bases; pH	Molarity; pH calculations; neutralization
8	Thermochemistry; specific heat; Hess's Law	$q = mc\Delta T$; Hess's Law problems
9	Reaction rates; equilibrium; Le Chatelier's Principle	Writing K expressions; predicting equilibrium shifts
10	Redox; electrochemistry; organic chemistry; nuclear chemistry	Oxidation numbers; cell voltage; half-life

Physics Topics Previewed in This Course

This table maps chemistry concepts to the physics topics they prepare you for:

Chemistry Topic	Physics Topic
States of matter; KMT	Thermodynamics; statistical mechanics
Atomic structure	Quantum mechanics; nuclear physics
Chemical bonding (ionic, covalent, metallic)	Electromagnetism; solid-state physics
The mole; Avogadro's number	Statistical mechanics; Boltzmann constant
Stoichiometry; conservation of mass	Conservation laws
Gas laws; PV = nRT	Thermodynamics; ideal gas behavior in engines
Acids/bases; pH (logarithms)	Logarithmic scales in physics (decibels, magnitudes)
Specific heat; thermochemistry	Thermal physics; First Law of Thermodynamics
Activation energy; reaction rates	Energy barriers; differential equations
Electrochemistry; cell voltage	Electrostatics; circuits; Ohm's Law
Nuclear reactions; half-life	Nuclear physics; radioactivity
$E = mc^2$	Special relativity; mass-energy equivalence

Final Word

Chemistry is sometimes described as the "central science" because it sits at the intersection of physics (below) and biology (above). The atoms you studied in Chapter 2 are the building blocks of matter that physics describes with forces and fields. Those same atoms, bonded together in the molecules of Chapter 3, form the proteins, DNA, and metabolic machinery that biology studies.

The ten chapters of this course have given you the vocabulary, the quantitative skills, and the conceptual framework to enter physics with confidence. You know what matter is made of, how it changes, how it stores and releases energy, and how its behavior scales from individual atoms to the bulk materials we encounter every day. Physics will take these foundations and ask the deeper questions: *why* do atoms behave as they do, *what forces* govern their interactions, and *how does energy flow* through the universe.

The boundary between chemistry and physics has never been sharp, and the best scientists have always worked on both sides of it. Welcome to physical science.