ANSWER KEY

General Chemistry 1: Practice Exam

Quick answers first, then a worked solution for every problem.

#	Topic	Answer
1	Percent composition	35.0% N
2	Empirical and molecular formula	CH_2O; $C_6H_{12}O_6$
3	Limiting reagent and percent yield	Cl_2 is limiting; 80.0% yield
4	Dilution	20.8 mL
5	Titration with a diprotic acid	0.06480 M
6	Net ionic equation	$Pb^{2+}(aq) + 2\ I^-(aq) \rightarrow PbI_2(s)$; spectators K^+ and NO_3^-
7	Oxidation numbers and redox	Fe reduced, C oxidized; Fe_2O_3 oxidizing agent, CO reducing agent
8	Ideal gas law	1.46 L
9	Calorimetry	0.454 J/(g·°C)
10	Hess's law	−110.5 kJ
11	Enthalpy of formation	−890.3 kJ
12	Electron configurations and quantum numbers	Fe [Ar] $4s^2\ 3d^6$; Fe^{3+} [Ar] $3d^5$; (2, 2, 0, +½) not allowed
13	Periodic trends	(a) Na > Mg > S > Cl (b) Na < Al < Mg
14	Lewis structures, resonance and formal charge	3 resonance structures; N +1, O 0/−1/−1; bond order 4/3
15	VSEPR, polarity, hybridization and intermolecular forces	NH_3 trigonal pyramidal, polar, sp^3; BF_3 trigonal planar, nonpolar, sp^2; Ne < CH_4 < H_2S < H_2O

1. Percent composition

What is the percent nitrogen by mass in ammonium nitrate, NH_4NO_3?

1. Molar mass: 2(14.01) + 4(1.008) + 3(16.00) = 80.05 g/mol.
2. Mass of N in one mole: 2 × 14.01 = 28.02 g.
3. 28.02 / 80.05 × 100 = **35.0% N**.

> **How to explain it:** Count every atom of the element, not just one. Students miss that NH_4NO_3 has two nitrogens, one in each ion, and get 17.5%.

2. Empirical and molecular formula

A compound is 40.00% C, 6.71% H and 53.29% O by mass. Its molar mass is 180.16 g/mol. Find the empirical and molecular formulas.

1. Assume 100 g: 40.00 g C, 6.71 g H, 53.29 g O.
2. Moles: C = 40.00/12.01 = 3.331; H = 6.71/1.008 = 6.657; O = 53.29/16.00 = 3.331.
3. Divide by the smallest (3.331): C 1, H 2, O 1. Empirical formula **CH_2O** (30.03 g/mol).
4. 180.16 / 30.03 = 6, so the molecular formula is **$C_6H_{12}O_6$** (glucose).

> **How to explain it:** "Percent to mass, mass to mole, divide by small, multiply till whole." The common mistake is dividing the percentages by each other instead of converting to moles first.

3. Limiting reagent and percent yield

$2\ Al + 3\ Cl_2 \rightarrow 2\ AlCl_3$. You react 10.0 g Al with 30.0 g Cl_2 and collect 30.1 g $AlCl_3$. Which reactant is limiting, and what is the percent yield?

1. Moles: Al = 10.0/26.98 = 0.3706 mol; Cl_2 = 30.0/70.90 = 0.4231 mol.
2. Al needs 0.3706 × 3/2 = 0.5560 mol Cl_2, but only 0.4231 mol is there, so **Cl_2 is limiting**.
3. Theoretical $AlCl_3$: 0.4231 × 2/3 = 0.2821 mol × 133.33 g/mol = 37.61 g.
4. Percent yield = 30.1 / 37.61 × 100 = **80.0%**.

> **How to explain it:** The reactant with the larger mass is limiting here. That's the whole lesson: compare moles through the coefficients, never grams. Using Al by mistake gives a theoretical yield of 49.4 g.

4. Dilution

How many mL of 6.0 M HCl do you need to make 250. mL of 0.50 M HCl?

1. $M_1V_1 = M_2V_2$, so $V_1 = (0.50 \times 250.) / 6.0 =$ **20.8 mL**.
2. Then add water until the total volume is 250. mL.

> **How to explain it:** Moles of HCl don't change when you add water, only the volume does. Students sometimes add 250 mL of water to the 20.8 mL instead of diluting up to 250 mL total.

5. Titration with a diprotic acid

25.00 mL of H_2SO_4 is neutralized by 32.40 mL of 0.1000 M NaOH. What is the H_2SO_4 concentration?

1. $H_2SO_4 + 2\ NaOH \rightarrow Na_2SO_4 + 2\ H_2O$.
2. Moles NaOH = $0.03240\ L \times 0.1000\ M = 3.240 \times 10^{-3}$ mol.
3. Moles $H_2SO_4 = 3.240 \times 10^{-3} / 2 = 1.620 \times 10^{-3}$ mol.
4. $[H_2SO_4] = 1.620 \times 10^{-3} / 0.02500\ L =$ **0.06480 M**.

> **How to explain it:** $M_1V_1 = M_2V_2$ only works for 1:1 reactions. Using it here gives 0.1296 M, exactly double.

6. Net ionic equation

Write the net ionic equation for mixing $Pb(NO_3)_2(aq)$ with $KI(aq)$. Name the spectator ions.

1. Molecular: $Pb(NO_3)_2(aq) + 2\ KI(aq) \rightarrow PbI_2(s) + 2\ KNO_3(aq)$.
2. Split the aqueous strong electrolytes into ions; keep the solid together.
3. Cancel what appears on both sides: **$Pb^{2+}(aq) + 2\ I^-(aq) \rightarrow PbI_2(s)$**. Spectators: K^+ and NO_3^-.

> **How to explain it:** Only the ions that actually change partners stay. Nitrates and Group 1 salts are always soluble, so they're always spectators. Don't split the precipitate into ions.

7. Oxidation numbers and redox

$Fe_2O_3 + 3\ CO \rightarrow 2\ Fe + 3\ CO_2$. What is oxidized, what is reduced, and which are the oxidizing and reducing agents?

1. Fe goes from +3 in Fe_2O_3 to 0 in Fe: gains electrons, so **Fe is reduced**.
2. C goes from +2 in CO to +4 in CO_2: loses electrons, so **C is oxidized**.
3. **Fe_2O_3 is the oxidizing agent** (it gets reduced). **CO is the reducing agent** (it gets oxidized).

> **How to explain it:** OIL RIG (oxidation is loss, reduction is gain). The agent is the opposite of what happens to it, and the agent is the whole compound, not just the atom.

8. Ideal gas law

What volume does 2.50 g of CO_2 occupy at 25.0 °C and 0.950 atm?

1. n = 2.50 / 44.01 = 0.05681 mol.
2. T = 25.0 + 273.15 = 298.15 K.
3. V = nRT/P = (0.05681)(0.08206)(298.15) / 0.950 = **1.46 L**.

> **How to explain it:** Always Kelvin. Plugging in 25 °C gives 0.123 L, which should look suspiciously small for over 2 grams of gas.

9. Calorimetry

A 45.0 g piece of metal at 95.0 °C is dropped into 100.0 g of water at 22.0 °C. The final temperature is 25.4 °C. Find the metal's specific heat.

1. Heat gained by water: q = (100.0)(4.184)(25.4 − 22.0) = 1,423 J.
2. Heat lost by metal = −1,423 J, with ΔT = 25.4 − 95.0 = −69.6 °C.
3. c = 1,423 / (45.0 × 69.6) = **0.454 J/(g·°C)** (close to iron).

> **How to explain it:** Heat lost by the hot thing equals heat gained by the cold thing. Both end at the same final temperature, which is the step students forget.

10. Hess's law

Use these reactions to find ΔH for C(s) + ½ O_2(g) → CO(g).
C(s) + O_2(g) → CO_2(g), ΔH = −393.5 kJ
CO(g) + ½ O_2(g) → CO_2(g), ΔH = −283.0 kJ

1. Keep the first reaction as written (it has C on the left).
2. Reverse the second so CO is a product: CO_2 → CO + ½ O_2, ΔH = +283.0 kJ.
3. Add them: CO_2 cancels, and ΔH = −393.5 + 283.0 = **−110.5 kJ**.

> **How to explain it:** Match each species to where it sits in the target. Reversing flips the sign; multiplying multiplies ΔH. Forgetting the sign flip is the most common error.

11. Enthalpy of formation

Find ΔH° for CH_4(g) + 2 O_2(g) → CO_2(g) + 2 H_2O(l). ΔH°f: CH_4 −74.8, CO_2 −393.5, H_2O(l) −285.8 kJ/mol.

1. ΔH° = Σ products − Σ reactants.
2. Products: −393.5 + 2(−285.8) = −965.1 kJ. Reactants: −74.8 + 2(0) = −74.8 kJ.
3. ΔH° = −965.1 − (−74.8) = **−890.3 kJ**.

> **How to explain it:** Elements in their standard state (O_2) have ΔH°f = 0. Students forget to multiply water by 2, or subtract in the wrong order.

12. Electron configurations and quantum numbers

(a) Write the configurations of Fe and Fe^{3+}. (b) Which set of quantum numbers is not allowed: (2, 1, 0, +½), (3, 2, −2, −½), (2, 2, 0, +½)?

1. Fe: [Ar] $4s^2$ $3d^6$. Fe^{3+} removes the 4s electrons first, then one 3d: **[Ar] $3d^5$**.
2. l must be less than n, so **(2, 2, 0, +½) is not allowed**.

> **How to explain it:** Fill 4s before 3d, but empty 4s before 3d when making a cation. Writing Fe^{3+} as [Ar] $4s^2$ $3d^3$ is the classic mistake.

13. Periodic trends

(a) Rank Na, Mg, S and Cl by atomic radius. (b) Rank Na, Mg and Al by first ionization energy.

1. Radius shrinks across a period as nuclear charge grows: **Na > Mg > S > Cl**.
2. Ionization energy generally rises across, but Al's 3p electron is easier to remove than Mg's filled 3s: **Na < Al < Mg**.

> **How to explain it:** More protons pull the same shell in tighter. Then show the exception: a lone p electron sits higher in energy than a paired s electron, so Al breaks the trend. N and O are the other famous exception.

14. Lewis structures, resonance and formal charge

Draw the nitrate ion, NO_3^-. How many resonance structures are there, what are the formal charges, and what is the N–O bond order?

1. Valence electrons: 5 + 3(6) + 1 = 24.
2. N in the center with one N=O double bond and two N–O single bonds. There are **3 resonance structures**, one for each position of the double bond.
3. Formal charges: N = +1, double-bonded O = 0, each single-bonded O = -1. They add to -1, the ion's charge.
4. Bond order = 4 bonds shared over 3 positions = **4/3, about 1.33**. All three bonds are the same length.

> **How to explain it:** Formal charge = valence electrons − dots − lines. Resonance structures aren't flipping back and forth: the real ion is one blend of all three.

15. VSEPR, polarity, hybridization and intermolecular forces

(a) Give the electron geometry, shape, polarity and hybridization of NH_3 and BF_3. (b) Rank by boiling point: Ne, CH_4, H_2S, H_2O.

1. NH_3: 4 electron groups (3 bonds + 1 lone pair), tetrahedral electron geometry, **trigonal pyramidal**, **polar**, sp^3.
2. BF_3: 3 groups, no lone pairs, **trigonal planar**, **nonpolar** (the dipoles cancel), sp^2.
3. Boiling point: **Ne < CH_4 < H_2S < H_2O**. Ne and CH_4 have only dispersion forces (CH_4 is bigger), H_2S adds dipole–dipole, and H_2O has hydrogen bonding.

> **How to explain it:** Count electron groups first, then name the shape using only the atoms. Polar bonds don't make a polar molecule if the shape cancels them, and that's exactly the BF_3 trap.