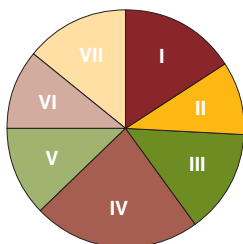


Chemistry: Content Knowledge (0245)

Test at a Glance

Test Name	Chemistry: Content Knowledge		
Test Code	0245		
Time	2 hours		
Number of Questions	100		
Format	Multiple-choice questions; calculator use prohibited		
	Content Categories	Approximate Number of Questions	Approximate Percentage of Examination
	I. Matter and Energy; Heat, Thermodynamics, and Thermochemistry	16	16%
	II. Atomic and Nuclear Structure	10	10%
	III. Nomenclature; the Mole, Chemical Bonding, and Geometry	14	14%
	IV. Periodicity and Reactivity; Chemical Reactions; Biochemistry and Organic Chemistry	23	23%
	V. Solutions and Solubility; Acid/Base Chemistry	12	12%
	VI. History and Nature of Science; Science, Technology, and Social Perspectives	11	11%
	VII. Mathematics, Measurement, and Data Management; Laboratory Procedures and Safety	14	14%



About This Test

The Chemistry: Content Knowledge test is designed to measure the knowledge and competencies necessary for a beginning teacher of secondary school Chemistry. Examinees have typically completed or nearly completed a bachelor's degree program with appropriate coursework in chemistry and education. This test may contain some questions that will not count towards your score.

The development of the test questions and the construction of the test reflect the National Science Education Standards (NSES) and the National Science Teacher Association (NSTA) standards and recognize that there are conceptual and procedural schemes that unify the various scientific disciplines. These fundamental concepts and processes (systems; models; constancy and change; equilibrium; form and function) are useful in understanding the natural world. Insofar as possible, then, the test questions will have the primary objective of evaluating the content areas by using questions that focus on conceptual understanding, critical thinking, and problem solving in science. The test content is developed and reviewed in collaboration with practicing high school chemistry teachers, teacher-educators, and higher education content specialists to keep the test updated and representative of current standards.

The 100 multiple choice questions include concepts, terms, phenomena, methods, applications, data analysis, and problem solving in Chemistry, and include an understanding of the impact of science and technology on the environment and human affairs. The topics are typically those covered in introductory college-level chemistry courses, although some questions of a more advanced nature are included, because secondary-school teachers must understand the subject matter from a more advanced viewpoint than that presented to their students.

Examinees will not need to use calculators in taking this test. The test book contains a periodic table of the elements and a table of information that presents various physical constants and a few conversion factors among SI units. Whenever necessary, additional values of physical constants are printed with the text of a question.

Topics Covered

Representative descriptions of topics covered in each category are provided below.

I. Matter and Energy; Heat and Thermodynamics, and Thermochemistry

A. Matter and Energy

1. Organization of matter
 - Pure substances (elements and compounds)
 - Mixtures (homogeneous, heterogeneous, solutions, suspensions)
 - States of matter (solid, liquid, gas, plasma)
2. Particulate structure of matter
 - Atoms, ions, molecules
3. Differences between chemical and physical properties and chemical and physical changes
 - Chemical vs. physical properties
 - Chemical vs. physical changes
 - Intensive vs. extensive properties
4. Conservation of energy and the conservation of matter in chemical processes
 - Law of conservation of energy
 - Law of conservation of matter
5. Different forms of energy
 - Kinetic and potential
 - Chemical, electrical, electromagnetic, nuclear, and thermal energy
 - Conversions between different forms of energy within chemical systems

B. Thermodynamics in Chemistry

1. Temperature, thermal energy, and heat capacity, including temperature scales, units of energy, and calculations involving these concepts
 - Temperature and temperature scales
 - Thermal energy and units of energy
 - Heat transfer
 - Heat capacity and specific heat
 - Calorimetry calculations
2. Concepts and calculations involving phase transitions between the various states of matter
 - Phase transitions and diagrams
 - Heats of vaporization, fusion, and sublimation
 - Heating curves
3. Kinetic molecular theory and ideal gas laws
 - Assumptions of the kinetic molecular theory
 - Ideal gases and the ideal gas laws (e.g., applications, calculations)
 - Real gas behavior
4. Energetics of chemical reactions
 - Exothermic and endothermic reactions
 - Bond energy; Hess's law
5. How the laws of thermodynamics relate to chemical reactions and phase changes
 - Laws of thermodynamics
 - Spontaneous/reversible processes
 - Change in enthalpy, entropy, and Gibbs energy in chemical/physical processes

II. Atomic and Nuclear Structure

1. Current model of atomic structure
 - Description of atomic model (e.g., subatomic particles, orbitals, quantum numbers)
 - Experimental basis (e.g., cathode ray tube, gold foil experiment, spectral lines)
 - Isotopes (mass number, average atomic mass)
2. Electron configuration of the elements based on the periodic table
 - Aufbau principle, Hund's rule, Pauli exclusion principle
 - Correlation between electron configuration and periodic table
 - Relationship between electron configuration and chemical and physical properties
3. Radioactivity
 - Characteristics of alpha particles, beta particles, and gamma radiation
 - Radioactive decay processes; half life
 - Fission, fusion, and other nuclear reactions
 - Balancing nuclear reactions and identifying products of nuclear reactions
4. How the electronic absorption and emission spectra of elements are related to electron energy levels
 - Electronic energy transitions in atoms (e.g., ground state, excited states, emission/absorption of energy)
 - Energy of electronic absorption/emission spectral lines in various regions of the electromagnetic spectrum
 - Relationship between energy, frequency, and wavelength

III. Nomenclature; the Mole, Chemical Bonding, and Geometry

A. Nomenclature and Chemical Composition

1. Systematic names and chemical formulas of simple inorganic compounds
 - Binary compounds
 - Acids, bases, and salts
 - Hydrates
2. Names of common organic compounds based on their functional groups
 - Alkanes, alkenes, and alkynes
 - Alcohols, ethers, ketones, aldehydes, amines
3. Mole concept and how it applies to chemical composition
 - Avogadro's number, molar mass, and mole conversions
 - Calculation of empirical and molecular formulas
 - Percent composition

B. Bonding and Structure

1. Common properties of bonds
 - Relative bond lengths
 - Relative bond strengths
2. Bond types
 - Ionic bonding
 - Covalent bonding (polar, nonpolar, hybridization)
 - Metallic bonding
3. Structural formulas and molecular geometry (shape)
 - Lewis structures including formal charges
 - Resonance structures
 - Molecular geometry (shape and approximate bond angles)
4. Identify polar and nonpolar molecules
 - Analysis of bonding in the molecule
 - Symmetry of molecular structure

5. Intermolecular interactions
 - Hydrogen bonding
 - London forces (instantaneous induced dipole-dipole)
 - Dipole-dipole
 - Dipole-induced dipole
6. How bonding and structure correlate with physical properties
 - Boiling points and melting points
 - Solubility
 - Equilibrium vapor pressure

IV. Periodicity and Reactivity; Chemical Reactions; Biochemistry and Organic Chemistry

A. Periodicity

1. Basis of the periodic table and general layout
 - Arranged in groups and periods
 - Atomic number and mass
 - Symbols of the elements
 - Metals, nonmetals, metalloids
 - Transition elements
2. Periodic trends in physical and chemical properties of the elements
 - Atomic/ionic radius
 - Ionization energy
 - Electron affinity
 - Electronegativity
 - Physical properties (e.g., boiling/melting points, conductivity)
 - Chemical reactivity

B. Chemical Reactions and Basic Principles

- Balancing chemical equations
 - Simple chemical equations
 - Chemical equations involving oxidation-reduction
- Stoichiometric calculations
 - Simple calculations based on balanced chemical equations involving moles, mass, and volume
 - Limiting reagent calculations and percent yield
- Identify, write, and predict products of simple reaction types
 - Combustion, neutralization
 - Decomposition, dehydration
 - Single and double replacement
 - Oxidation-reduction
- Chemical kinetics
 - Rate laws, rate constants, and reaction order
 - Activation energy and reaction mechanisms including catalysts
 - Factors affecting reaction rate such as concentration, surface area, and temperature
- Chemical reaction equilibrium
 - Equilibrium constants
 - Le Châtelier's principle
- Oxidation-reduction reactions and how to determine oxidation states
 - Oxidation states
 - Identify oxidation-reduction reactions and half reactions
 - Standard reduction potentials
 - Electrochemical reactivity series
 - Electrochemical cells

C. Biochemistry and Organic Chemistry

- Important biochemical compounds
 - Carbohydrates, including simple sugars
 - Lipids
 - Proteins and amino acids
 - DNA and RNA
 - Products of photosynthesis and respiration
- Common organic compounds (i.e., identify functional groups)
 - Alcohols
 - Ketones and aldehydes
 - Alkanes, alkenes, and alkynes
 - Ethers
 - Carboxylic acids
 - Amines
 - Benzene

V. Solutions and Solubility; Acid-Base Chemistry

A. Solutions and Solubility

- Solution terminology and calculations
 - Dilute, concentrated
 - Saturated, unsaturated, supersaturated
 - Solvent, solute
 - Concentration units (e.g., molarity, molality, mole fraction, parts per million (ppm), parts per billion (ppb), percent by mass or volume)
 - Preparation of solutions of varying concentrations
- Factors affecting solubility and dissolution rate
 - Dissolution rate (i.e., temperature, pressure, surface area, agitation)
 - Solubility and solubility curves (temperature and pressure dependent)
- Solution phenomena based on colligative properties
 - Freezing point depression
 - Boiling point elevation
 - Vapor pressure effects
 - Osmotic pressure

4. Common applications of equilibrium in ionic solutions
 - Solubility of ionic compounds (e.g., solubility rules, slightly soluble compounds)
 - K_{sp} calculations including percent dissociation and precipitation
 - Common ion effect
 - Electrolytes, nonelectrolytes, and electrical conductivity

B. Acid-Base Chemistry

1. Define and identify acids and bases and know their properties
 - Arrhenius acids and bases
 - Brønsted-Lowry acids and bases
 - Lewis acids and bases
 - Neutralization and equivalence point
2. The pH scale and calculations involving pH and pOH
 - pH scale
 - Calculation of pH and pOH
 - Calculation of $[H^+]$ and $[OH^-]$
 - Knows the meaning of K_w
3. Concepts and calculations involving acid-base titrations
 - Use and selection of indicators (e.g., phenolphthalein, litmus paper)
 - Endpoint determination
 - Calculations based on titrations
4. Equilibrium relationships in acid-base chemistry
 - Strong/weak acids and bases, including common examples
 - Monoprotic and polyprotic acids
 - K_a , K_b , and percent dissociation
 - Hydrolysis (acidic and basic salts)
 - Buffer solutions

VI. History and Nature of Science; Science, Technology, and Social Perspectives

A. History and Nature of Scientific Inquiry

1. Processes involved in scientific inquiry
 - Formulating problems
 - Forming and testing hypotheses
 - Development of theories, models, and laws (postulates, assumptions)
 - Process skills including observing, concluding, comparing, inferring, categorizing, and generalizing
2. Experimental design
 - Testing hypotheses
 - Significance of controls
 - Use and identification of variables
 - Data collection planning
3. Nature of scientific knowledge
 - Subject to change
 - Consistent with experimental evidence
 - Reproducibility
 - Unifying concepts and processes (e.g., systems, models, constancy and change, equilibrium, form and function)
4. Major historical developments in chemistry and the contributions of major historical figures
 - How current chemical principles and models developed over time
 - Major developments in chemistry (e.g., atomic model, ideal gas behavior) including major historical figures

B. Science, Technology, Society, and the Environment

1. Impact of chemistry and technology on society and the environment
 - Pharmaceuticals
 - Acid rain
 - Medical imaging
 - Air and water pollution
 - Greenhouse gases
 - Ozone layer depletion
 - Waste disposal and recycling
 - Nanotechnology
2. Applications of chemistry in daily life
 - Plastics, soap, batteries, fuel cells, and other consumer products
 - Water purification
 - Chemical properties of household products
3. Advantages and disadvantages associated with various types of energy production
 - Renewable and nonrenewable energy resources
 - Conservation and recycling
 - Pros and cons of power generation based on various sources such as fossil and nuclear fuel, hydropower, wind power, solar power, and geothermal power

VII. Mathematics, Measurement, and Data Management; Laboratory Procedures and Safety

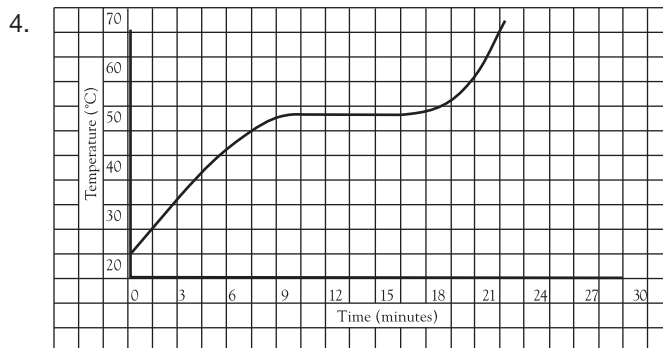
1. Collect, evaluate, manipulate, interpret, and report data
 - Significant figures in collected data and calculations
 - Organization and presentation of data
 - Knows how to interpret and draw conclusions from data presented in tables, graphs, and charts (e.g., trends in data, relationships between variables, predictions and conclusions based on data)
2. Units of measurement, notation systems, conversions, and mathematics used in chemistry
 - Standard units of measurement
 - Unit conversion
 - Scientific notation
 - Measurement equipment
3. Basic error analysis
 - Determining mean
 - Accuracy and precision
 - Identifying sources and effects of error
 - Percent error
4. Appropriate preparation, use, storage, and disposal of materials in the laboratory
 - Appropriate use and storage
 - Safe disposal
 - Preparation for classroom use
 - Safe procedures and safety precautions
5. Appropriate use, maintenance, and calibration of laboratory equipment
 - Appropriate use and storage
 - Maintenance and calibration
 - Preparation for classroom use
 - Safety procedures and precautions when using equipment
6. Safety procedures and precautions for the high school chemistry laboratory
 - Location and use of standard safety equipment such as eyewash and shower
 - Laboratory safety rules for students
 - Appropriate apparel and conduct in the laboratory, such as wearing goggles
 - Emergency procedures

Sample Test Questions

The sample questions that follow illustrate the kinds of questions in the test. They are not, however, representative of the entire scope of the test in either content or difficulty. Answers with explanations follow the questions.

Directions: Each of the sentences or incomplete statements below is followed by four suggested answers or completions. Select the one that is best in each case.

- In a laboratory experiment, crystals are heated in a dry glass test tube using a Bunsen burner. During heating, a clear liquid is observed inside the mouth of the test tube. Which of the following is the most reasonable conclusion drawn from this observation?
 - The gas fuel used to heat the crystals forms water as it burns.
 - The crystals give off water when heated.
 - The crystals give off both hydrogen and oxygen gases that combine to form water.
 - Condensation from the air collects on the test tube as the crystals are heated.
- Suppose that a mixture of 8 g of sugar, 5.20 g of salt, and 100.01 g of flour is prepared. What is the total mass of the mixture expressed in exponential notation and with the correct number of significant figures?
 - 1×10^2
 - 1.1×10^2
 - 1.13×10^2
 - 1.132×10^2
- Use of a small quantity of which of the following gases in a classroom requires special consideration because the gas is poisonous?
 - Steam
 - Hydrogen
 - Hydrogen sulfide
 - Carbon dioxide



A sample of a pure solid substance is heated at a constant rate and its temperature recorded as a function of time. A graph of the data is shown above. At about what temperature is the heat added being used to melt the substance?

- 25°C
 - 41°C
 - 53°C
 - 60°C
- Changes in which the entropy of the system increases include which of the following?
 - Melting ice at room temperature
 - Evaporating water at room temperature
 - Dissolving NaCl in room-temperature water
 - I only
 - III only
 - I and II only
 - I, II, and III

6.

Isotope	Isotopic Mass (amu)	Percent Abundance
41	40.9	10.0%
44	43.9	30.0%
46	45.9	60.0%

A fictional element with the three naturally occurring isotopes described above would be listed in the periodic table with an atomic mass of

- (A) 42.1
 (B) 43.6
 (C) 44.8
 (D) 45.9

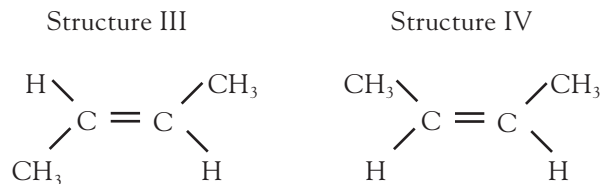
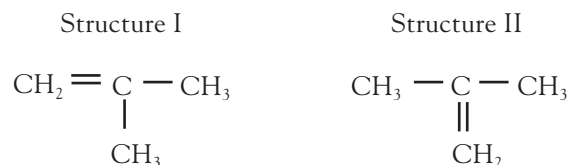
7. In an attempt to compare the half-lives of two radioactive elements, X and Y, a scientist set aside 400 g of each. After six months, the scientist found that 25 g of X and 200 g of Y remained. Which of the following statements is true?

- (A) The half-life of Y is twice the half-life of X.
 (B) The half-life of Y is four times the half-life of X.
 (C) The half-life of Y is eight times the half-life of X.
 (D) Unless the exact time interval is established, a comparison cannot be made.

8. Which of the following statements is correct about any chemical reaction that is at equilibrium?

- (A) The molecules stop reacting.
 (B) Only side reactions continue; the main reaction stops.
 (C) Forward and backward reactions occur at equal rates.
 (D) There are as many molecules of reactant as there are molecules of product.

9. Which, if any, of the following structural formulas represent the same compound?



- (A) I and II only
 (B) III and IV only
 (C) I, II, III, and IV
 (D) None of the formulas represent identical compounds.
10. The correct formula for copper (I) sulfate is
- (A) CuSO_4
 (B) Cu_2SO_4
 (C) Cu_4SO
 (D) Cu_4SO_4
11. The pH of a 4.0×10^{-4} M HCl solution is between
- (A) 2 and 3
 (B) 3 and 4
 (C) 4 and 5
 (D) 5 and 6

12. $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag(s)}$ $E^{\circ} = +0.80 \text{ V}$
 $\text{Cr}^{3+} + 3\text{e}^- \rightarrow \text{Cr(s)}$ $E^{\circ} = -0.74 \text{ V}$
- Based on the standard reduction potentials for chromium and silver shown above, what is the cell potential for the reaction below?
- $$3 \text{Ag}^+ + \text{Cr(s)} \rightarrow 3 \text{Ag(s)} + \text{Cr}^{3+}$$
- (A) 0.06 V
(B) 1.54 V
(C) 1.66 V
(D) 3.14 V
13. Which of the following is an important function of carbohydrates in living organisms?
- (A) They are the primary component of enzymes.
(B) They constitute a source of energy for the organism.
(C) They contain the genetic information of the cell.
(D) They comprise structural parts such as cartilage and tendons.
14. Liquids with molecules held together by van der Waals forces have which of the following properties?
- (A) High solubilities in water
(B) High melting points
(C) Low boiling points
(D) Significant electrical conductivities in the solid phase
15. The solubility product, K_{sp} , for Mg(OH)_2 is 1.0×10^{-11} . What is the concentration of Mg^{2+} in a saturated solution of this base?
- (A) $\sqrt{5.0 \times 10^{-12}} \text{ M}$
(B) $\sqrt{1.0 \times 10^{-11}} \text{ M}$
(C) $\sqrt[3]{2.5 \times 10^{-12}} \text{ M}$
(D) $\sqrt[3]{1.0 \times 10^{-11}} \text{ M}$
16. When 0.50 mol of octane, C_8H_{18} , is burned completely and the reaction products are brought to 20°C and 1 atmosphere, the products include approximately
- (A) 18 moles of water
(B) Close to 100 liters of carbon dioxide
(C) Close to 180 liters of carbon dioxide
(D) Close to 200 liters of water vapor
17. Which of the following properties of a substance depends on the amount of the sample?
- (A) Temperature
(B) Half-life
(C) Density
(D) Inertia
18. What quantity of oxygen, O_2 , contains very nearly the same number of molecules as 36.0 grams of water, H_2O ?
- (A) 64.0 grams
(B) 32.0 grams
(C) 16.0 grams
(D) 8.0 grams
19. In the reaction of solutions of silver nitrate and sodium phosphate, a silver precipitate is formed. The balanced ionic equation for this reaction is
- (A) $\text{Ag}^+ + \text{PO}_4^{3-} \rightleftharpoons \text{Na}^+ + \text{AgPO}_4(\text{s})$
(B) $\text{Ag}^+ + \text{PO}_4^{3-} \rightleftharpoons \text{AgPO}_4(\text{s})$
(C) $3 \text{Ag}^+ + \text{PO}_4^{3-} \rightleftharpoons \text{Ag}_3\text{PO}_4(\text{s})$
(D) $4 \text{Ag}^+ + \text{PO}_4^{3-} \rightleftharpoons \text{Ag}_4\text{PO}_4(\text{s})$
20. The two elements in which of the following pairs have a valence of +3 ?
- (A) Al and Ga
(B) N and O
(C) Li and Na
(D) F and Cl

Answers

1. The crystals when heated may give off water in the form of steam. When this moist air reaches the top of the tube, condensation occurs inside the top of the tube because the tube is cooler than the rising warmer air. The correct answer is B.

2. For addition, the correct number of significant figures to the right of the decimal point in the answer is the same as the smallest number of digits to the right of the decimal point in any of the terms to be added. In this example, that number is zero, which gives 113. as the total mass of the mixture to the correct number of significant figures. The number is then converted to exponential notation. The correct answer is C.

3. Of the gases listed, only hydrogen sulfide is poisonous in small quantities. The correct answer is C.

4. When a substance is heated, its temperature increases unless it is undergoing a phase change. During melting, the temperature remains constant since the energy absorbed is being used to do work against the attractive forces in becoming liquid particles. In the diagram, melting begins around 9 minutes and a temperature around 53°C. The correct answer is C.

5. Entropy is a measure of disorder. In all three cases, the disorder of the system increases. The correct answer is D.

6. A quick calculation using the isotope number, $(0.1 \times 40.9) + (0.3 \times 43.9) + (0.6 \times 45.9) = 44.8$, gives the correct answer, C.

7. Element X decayed from 400 g to 25 g, in a time period of 4 half-lives. Element Y decayed from 400 g to 200 g, in a time period of 1 half-life. Y decays slower than X, and the half-life of Y is 4 times that of X. The correct answer is B.

8. The correct answer is C. The definition of an equilibrium is that the forward and backward reactions occur at equal rates.

9. The correct answer is A. I and II are the same compound, 2-methyl-1-propene. III and IV are *cis* and *trans* isomers. They are geometric isomers with different properties.

10. The correct answer is B. Copper (I) is Cu^+ and sulfate is SO_4^{2-} ; therefore, for charge neutrality the compound is Cu_2SO_4 .

11. Choice B is the correct answer. HCl dissociates completely, $[\text{H}^+] = 4.0 \times 10^{-4} \text{ M}$ and the pH is between 3 and 4.

12. The correct answer is B.

$$E = E^0(\text{Ag}) - E^0(\text{Cr}) = 0.80 + 0.74 = 1.54$$

13. Choice B is the correct answer. Carbohydrates are broken down to glucose during digestion, and glucose is oxidized in the cell to produce energy in the form of ATP.

14. Van der Waals forces is the collective name for weak attractive forces between molecules. In general, liquids held together only by these forces have low boiling points relative to their molecular weights since only weak forces must be overcome during vaporization. Thus, C is the correct answer.

15. The K_{sp} of a salt is the product of the ion concentrations in a saturated solution. In the present case, $K_{sp} = [\text{Mg}^{2+}][\text{OH}^-]^2$. Since $[\text{OH}^-] = 2[\text{Mg}^{2+}]$, $K_{sp} = [\text{Mg}^{2+}](2[\text{Mg}^{2+}])^2 = 4[\text{Mg}^{2+}]^3 = 1.0 \times 10^{-11}$. Solving for Mg^{2+} one obtains $[\text{Mg}^{2+}] = [1.0 \times 10^{-11} / 4]^{1/3}$.

The correct answer is C.

16. The equation for the reaction is $2 \text{C}_8\text{H}_{18} + 25 \text{O}_2 \rightarrow 16 \text{CO}_2 + 18 \text{H}_2\text{O}$.

Choice B is correct; 0.5 mol octane produces 4 mol of CO_2 , which, at 20°C, occupies

$$\frac{(4 \text{ mol} \times 22.4 \text{ L/mol}) 293 \text{ K}}{273 \text{ K}} = 96 \text{ L} \cong 100 \text{ L}$$

17. Inertia is a property of a substance, proportional to its mass, and therefore depends on the amount of the sample. The correct answer is D.

18. The correct answer is A. 36 grams of water is 2 moles (2×18 grams). A 2-mole sample of O_2 contains the same number of molecules as does 2 moles of any other substance. A 2-mole sample of O_2 would have a mass of 2×32.0 grams = 64.0 grams.

19. The correct answer is C. The precipitate formed is Ag_3PO_4 . Atoms and net charge must be conserved in a chemical reaction, so the reactants and products must have equal numbers of each type of atom and the same net charge.

20. The correct answer is A. The valences of Al and Ga are both +3.



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