

Chemistry I

Fall Semester Objectives

Chapter 1 Matter and Change

1. Define *chemistry*.
2. Distinguish between the physical properties and chemical properties of matter.
3. Classify changes of matter as physical or chemical.
4. Explain the gas, liquid, and solid states in terms of particles.
5. Use a periodic table to name elements, given their symbols.
6. Use a periodic table to write the symbols of elements, given their names.
7. Describe the arrangement of the periodic table.
8. List the characteristics that distinguish metals, nonmetals, and metalloids.

Chapter 2 Measurements and Calculations

9. Describe the purpose of the scientific method.
10. Distinguish between qualitative and quantitative observations.
11. Distinguish between a quantity, a unit, and a measurement standard.
12. Name SI units for length, mass, time, volume, and density.
13. Distinguish between mass and weight.
14. Perform density calculations.
15. Transform a statement of equality to a conversion factor.
16. Distinguish between accuracy and precision.
17. Determine the number of significant figures in measurements.
18. Perform mathematical operations involving significant figures.
19. Convert measurements into scientific notation.
20. Distinguish between inversely and directly proportional relationships.

Chapter 3 Atoms: The Building Blocks of Matter

21. Explain the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.
22. Summarize the five essential points of Dalton's atomic theory.
23. Explain the relationship between Dalton's atomic theory and the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.
24. Summarize the observed properties of cathode rays that led to the discovery of the electron.
25. Summarize the experiment carried out by Rutherford and his co-workers that led to the discovery of the nucleus.
26. List the properties of protons, neutrons, and electrons.
27. Define *atom*.
28. Explain what isotopes are.
29. Define *atomic number* and *mass number*, and describe how they apply to isotopes.
30. Given the identity of a nuclide, determine its number of protons, neutrons, and electrons.
31. Define *mole* in terms of Avogadro's number, and define *molar mass*.
32. Solve problems involving mass in grams, amount in moles, and number of atoms of an element.

Chapter 4 Arrangement of Electrons in Atoms

33. Explain the mathematical relationship between the speed, wavelength, and frequency of electromagnetic radiation.

34. Discuss the dual wave-particle nature of light.
35. Discuss the significance of the photoelectric effect and the line-emission spectrum of hydrogen to the development of the atomic model.
36. Describe the Bohr model of the hydrogen atom.
37. Discuss Louis de Broglie's role in the development of the quantum model of the atom.
38. Compare and contrast the Bohr model and the quantum model of the atom.
39. Explain how the Heisenberg uncertainty principle and the Schrödinger wave equation led to the idea of atomic orbitals.
40. List the four quantum numbers, and describe their significance.
41. Relate the number of sublevels corresponding to each of an atom's main energy levels, the number of orbitals per sublevel, and the number of orbitals per main energy level.
42. List the total number of electrons needed to fully occupy each main energy level.
43. State the Aufbau principle, the Pauli exclusion principle, and Hund's rule.
44. Describe the electron configurations for the atoms of any element using orbital notation, electron-configuration notation, and, when appropriate, noble-gas notation.

Chapter 5 The Periodic Law

45. Explain the roles of Mendeleev and Moseley in the development of the periodic table.
46. Describe the modern periodic table.
47. Explain how the periodic law can be used to predict the physical and chemical properties of elements.
48. Describe how the elements belonging to a group of the periodic table are interrelated in terms of atomic number.
49. Describe the relationship between electrons in sublevels and the length of each period of the periodic table.
50. Locate and name the four blocks of the periodic table. Explain the reasons for these names.
51. Discuss the relationship between group configurations and group numbers.
52. Describe the locations in the periodic table and the general properties of the alkali metals, the alkaline-earth metals, the halogens, and the noble gases.
53. Define *atomic* and *ionic radii*, *ionization energy*, *electron affinity*, and *electronegativity*.
54. Compare the periodic trends of atomic radii, ionization energy, and electronegativity, and state the reasons for these variations.
55. Define *valence electrons*, and state how many are present in atoms of each main-group element.
56. Compare the atomic radii, ionization energies, and electronegativities of the *d*-block elements with those of the main-group elements.

Chapter 6: Chemical Bonding

57. Define *chemical bond*.
58. Explain why most atoms form chemical bonds.
59. Describe ionic and covalent bonding.
60. Explain why most chemical bonding is neither purely ionic nor purely covalent.
61. Classify bonding type according to electronegativity differences.
62. Define *molecule* and *molecular formula*.
63. Explain the relationships between potential energy, distance between approaching atoms, bond length, and bond energy.
64. State the octet rule.
65. List the six basic steps used in writing Lewis structures.
66. Explain how to determine Lewis structures for molecules containing single bonds, multiple bonds, or both.
67. Explain why scientists use resonance structures to represent some molecules.
68. Compare and contrast a chemical formula for a molecular compound with one for an ionic compound.
69. Discuss the arrangements of ions in crystals.
70. Define *lattice energy* and explain its significance.

71. List and compare the distinctive properties of ionic and molecular compounds.
72. Write the Lewis structure for a polyatomic ion given the identity of the atoms combined and other appropriate information.
73. Describe the electron-sea model of metallic bonding, and explain why metals are good electrical conductors.
74. Explain why metal surfaces are shiny.
75. Explain why metals are malleable and ductile but ionic-crystalline compounds are not.
76. Explain VSEPR theory.
77. Predict the shapes of molecules or polyatomic ions using VSEPR theory.
78. Explain how the shapes of molecules are accounted for by hybridization theory.
79. Describe dipole-dipole forces, hydrogen bonding, induced dipoles, and London dispersion forces.
80. Explain what determines molecular polarity.

Chapter 7 Chemical Formulas and Chemical Compounds

81. Explain the significance of a chemical formula.
82. Determine the formula of an ionic compound formed between two given ions.
83. Name an ionic compound given its formula.
84. Using prefixes, name a binary molecular compound from its formula.
85. Write the formula of a binary molecular compound given its name.
86. Give the oxidation number for each element in the formula of a chemical compound.
87. Calculate the formula mass or molar mass of any given compound.
88. Use molar mass to convert between mass in grams and amount in moles of a chemical compound.
89. Calculate the number of molecules, formula units, or ions in a given molar amount of a chemical compound.
90. Calculate the percentage composition of a given chemical compound.
91. Define *empirical formula*, and explain how the term applies to ionic and molecular compounds.
92. Determine an empirical formula from either a percentage or a mass composition.
93. Explain the relationship between the empirical formula and the molecular formula of a given compound.
94. Determine a molecular formula from an empirical formula.

Chapter 8 Chemical Equations and Reactions

95. List three observations that suggest that a chemical reaction has taken place.
96. List three requirements for a correctly written chemical equation.
97. Write a word equation and a formula equation for a given chemical reaction.
98. Balance a formula equation by inspection.
99. Define and give general equations for *synthesis*, *decomposition*, *single-replacement*, and *double-replacement* reactions.
100. Classify a reaction as synthesis, decomposition, single-replacement, double-replacement, or combustion.
101. List three types of synthesis reactions and six types of decomposition reaction.
102. List four types of single-replacement reactions and three types of double-replacement reactions.
103. Predict the products of simple reactions given the reactants.
104. Explain the significance of an activity series.
105. Use an activity series to predict whether a given reaction will occur and what the products will be.

Chapter 9 Stoichiometry

106. Define *stoichiometry*.
107. Describe the importance of the *mole ratio* in stoichiometric calculations.
108. Write a mole ratio relating two substances in a chemical equation.

109. Calculate the amount in moles of a reactant or product from the amount in moles of a different reactant or product.
110. Calculate the mass of a reactant or product from the amount in moles of a different reactant or product.
111. Calculate the amount in moles of a reactant or product from the mass of a different reactant or product.
112. Calculate the mass of a reactant or product from the mass of a different reactant or product.
113. Describe a method for determining which of two reactants is a limiting reactant.
114. Calculate the amount in moles or mass in grams of a product, given the amounts in moles or masses in grams of two reactants, one of which is in excess.
115. Distinguish between theoretical yield, actual yield, and percent yield.
116. Calculate percent yield, given the actual yield and quantity of a reactant.
117. Write equations for the dissolution of soluble ionic compounds in water.
118. Predict whether a precipitate will form when solutions of soluble ionic compounds are combined, and write net ionic equations for precipitation reactions.
119. Define *standard molar volume of a gas*, and use it to calculate gas masses and volumes.
120. Use standard molar volume to calculate the molar mass of a gas.
121. Use a chemical equation to specify volume ratios for gaseous reactants or products, or both.
122. Use volume ratios and the gas laws to calculate volumes, masses, or molar amounts of gaseous reactants or products.

Chemistry I

Spring Semester Objectives

Chapter 10: Physical Characteristics of Gases

123. State the kinetic-molecular theory of matter, and describe how it explains certain properties of matter.
124. List the five assumptions of the kinetic-molecular theory of gases. Define the terms *ideal gas* and *real gas*.
125. Describe each of the following characteristic properties of gases: expansion, density, fluidity, compressibility, diffusion, and effusion.
126. Describe the conditions under which a real gas deviates from “ideal” behavior.
127. Define *pressure*, and relate it to force.
128. Describe how pressure is measured.
129. Convert units of pressure.
130. State the standard conditions of temperature and pressure.
131. Use the kinetic-molecular theory to explain the relationships between gas volume, temperature, and pressure.
132. Use Boyle’s law to calculate volume-pressure changes at constant temperature.
133. Use Charles’s law to calculate volume-temperature changes at constant pressure.
134. Use Gay-Lussac’s law to calculate pressure-temperature changes at constant volume.
135. Use the combined gas law to calculate volume-temperature-pressure changes.
136. Use Dalton’s law of partial pressures to calculate partial pressures and total pressures.

Chapter 11: Molecular Composition of Gases

137. State the law of combining volumes.
138. State Avogadro’s law and explain its significance.
139. Define *standard molar volume of a gas*, and use it to calculate gas masses and volumes.
- 140.** Use standard molar volume to calculate the molar mass of a gas.
141. State the ideal gas law.
142. Derive the ideal gas constant and discuss its units.
143. Using the ideal gas law, calculate pressure, volume, temperature, or amount of gas when the other three quantities are known.
144. Using the ideal gas law, calculate the molar mass or density of a gas.
145. Reduce the ideal gas law to Boyle’s law, Charles’s law, and Avogadro’s law. Describe the conditions under which each applies.
146. Explain how Gay-Lussac’s law and Avogadro’s law apply to the volumes of gases in chemical reactions.
147. Use a chemical equation to specify volume ratios for gaseous reactants or products, or both.
148. Use volume ratios and the gas laws to calculate volumes, masses, or molar amounts of gaseous reactants or products.
149. State Graham’s law of effusion.
150. Determine the relative rates of effusion of two gases of known molar masses.
- 151.** State the relationship between the molecular velocities of two gases and their molar masses.

Chapter 12: Liquids and Solids

152. Describe the motion of particles in liquids and the properties of liquids according to the kinetic-molecular theory.
153. Discuss the process by which liquids can change into a gas. Define *vaporization*.
154. Discuss the process by which liquids can change into a solid. Define *freezing*.

155. Describe the motion of particles in solids and the properties of solids according to the kinetic-molecular theory.
156. Distinguish between the two types of solids.
157. Explain the relationship between equilibrium and changes of state.
158. Predict changes in equilibrium using Le Châtelier's principle.
159. Explain what is meant by equilibrium vapor pressure.
160. Describe the processes of boiling, freezing, melting, and sublimation.
161. Interpret phase diagrams.
162. Describe the structure of a water molecule.
163. Discuss the physical properties of water. Explain how they are determined by the structure of water.
164. Calculate the amount of heat energy absorbed or released when a quantity of water changes state.

Chapter 13: Solutions

165. Distinguish between a mixture and a pure substance.
166. Distinguish between heterogeneous and homogeneous mixtures.
167. List three different solute-solvent combinations.
168. Compare the properties of suspensions, colloids, and solutions.
169. Distinguish between electrolytes and nonelectrolytes.
170. List and explain three factors that affect the rate at which a solid solute dissolves in a liquid solvent.
171. Explain solution equilibrium, and distinguish among saturated, unsaturated, and supersaturated solutions.
172. Explain the meaning of "like dissolves like" in terms of polar and nonpolar substances.
173. List the three interactions that contribute to the heat of solution, and explain what causes dissolution to be exothermic or endothermic.
174. Compare the effects of temperature and pressure on solubility.
175. Given the mass of solute and volume of solvent, calculate the concentration of a solution.
176. Given the concentration of a solution, determine the amount of solute in a given amount of solution.
177. Given the concentration of a solution, determine the amount of solution that contains a given amount of solute.

Chapter 14: Ions in Aqueous Solutions and Colligative Properties

178. Write equations for the dissolution of soluble ionic compounds in water.
179. Predict whether a precipitate will form when solutions of soluble ionic compounds are combined, and write net ionic equations for precipitation reactions.
180. Compare dissociation of ionic compounds with ionization of molecular compounds.
181. Draw the structure of the hydronium ion, and explain why it is used to represent the hydrogen ion in solution.
182. Distinguish between strong electrolytes and weak electrolytes.
183. List four colligative properties, and explain why they are classified as colligative properties.
184. Calculate freezing-point depression, boiling-point elevation, and solution molality of nonelectrolytic solutions.
185. Calculate the expected changes in freezing point and boiling point of an electrolytic solution.
186. Discuss causes of the differences between expected and experimentally observed colligative properties of electrolytic solutions.

Chapter 15: Acids and Bases

187. List five general properties of aqueous acids and bases.
188. Name common binary acids and oxyacids, given their chemical formulas.
189. List five acids commonly used in industry and the laboratory, and give two properties of each.

190. Define *acid* and *base* according to Arrhenius's theory of ionization.
191. Explain the differences between strong and weak acids and bases.
192. Define and recognize *Brønsted acids* and *bases*.
193. Name compounds that are acids under the Lewis definition but are not acids under the Brønsted definition.
194. Describe a conjugate acid, a conjugate base, and an amphoteric compound.
195. Explain the process of neutralization.
196. Explain how acid rain damages marble structures.

Chapter 16: Acid-Base Titration and pH

197. Describe the self-ionization of water.
198. Define *pH*, and give the pH of a neutral solution at 25°C.
199. Explain and use the pH scale.
200. Given $[\text{H}_3\text{O}^+]$ or $[\text{OH}^-]$, find pH.
201. Given pH, find $[\text{H}_3\text{O}^+]$ or $[\text{OH}^-]$.
202. Describe how an acid-base indicator functions.
203. Explain how to carry out an acid-base titration.
204. Calculate the molarity of a solution from titration data.

Chapter 18: Chemical Equilibrium

205. Define *chemical equilibrium*.
206. Explain the nature of the equilibrium constant.
207. Write chemical equilibrium expressions and carry out calculations involving them.
208. Discuss the factors that disturb equilibrium.
209. Discuss conditions under which reactions go to completion.
210. Describe the common-ion effect.
211. Explain the concept of acid-ionization constants, and write acid-ionization equilibrium expressions.
212. Review the ionization constant of water.
213. Explain buffering.
214. Compare cation and anion hydrolysis.
215. Explain what is meant by *solubility-product constants*, and calculate their values.
216. Calculate solubilities using solubility-product constants.
217. Carry out calculations to predict whether precipitates will form when solutions are combined.

Chapter 17: Reaction Energy and Reaction Kinetics

218. Define *temperature*, and state the units in which it is measured.
219. Define *heat* and state its units.
220. Perform specific-heat calculations.
221. Explain heat of reaction, heat of formation, heat of combustion, and enthalpy change.
222. Solve problems involving heats of reaction, heats of formation, and heats of combustion.
223. Explain the relationship between enthalpy change and the tendency of a reaction to occur.
224. Explain the relationship between entropy change and the tendency of a reaction to occur.
225. Discuss the concept of free energy, and explain how the value of this quantity is calculated and interpreted.
226. Describe the use of free energy change to determine the tendency of a reaction to occur.
227. Explain the concept of reaction mechanism.
228. Use the collision theory to interpret chemical reactions.
229. Define *activated complex*.
230. Relate activation energy to heat of reaction.
231. Define *chemical kinetics*, and explain the two conditions necessary for chemical reactions to occur.

232. Discuss the five factors that influence reaction rate.
233. Define *catalyst*, and discuss two different types.
234. Explain and write rate laws for chemical reactions.

Chapter 19: Oxidation-Reduction Reactions

235. Assign oxidation numbers to reactant and product species.
236. Define *oxidation* and *reduction*.
237. Explain what an oxidation-reduction reaction (redox reaction) is.
238. Explain what must be conserved in redox equations.
239. Balance redox equations by using the half-reaction method.
240. Relate chemical activity to oxidizing and reducing strength.
241. Explain the concept of auto-oxidation.
242. Explain what is required for an electrochemical cell.
243. Describe the nature of voltaic cells.
244. Describe the nature of electrolytic cells.
245. Explain the process of electroplating.
246. Describe the chemistry of a rechargeable cell.
247. Calculate cell potentials from a table of standard electrode potentials

Chapter 22 Nuclear Chemistry

248. Explain what a nuclide is, and describe the different ways nuclides can be represented.
249. Define and relate the terms *mass defect* and *nuclear binding energy*.
250. Explain the relationship between nucleon number and stability of nuclei.
251. Explain why nuclear reactions occur, and know how to balance a nuclear equation.
252. Define and relate the terms *radioactive decay* and *nuclear radiation*.
253. Describe the different types of radioactive decay and their effects on the nucleus.
254. Define the term *half-life*, and explain how it relates to the stability of a nucleus.
255. Define and relate the terms *decay series*, *parent nuclide*, and *daughter nuclide*.
256. Explain how artificial radioactive nuclides are made, and discuss their significance.
257. Compare the penetrating ability and shielding requirements of alpha particles, beta particles, and gamma rays.
258. Define the terms *roentgen* and *rem*, and distinguish between them.
259. Describe three devices used in radiation detection.
260. Discuss applications of radioactive nuclides.
261. Define *nuclear fission*, *chain reaction*, and *nuclear fusion*, and distinguish between them.
262. Explain how a fission reaction is used to generate power.
263. Discuss the possible benefits and the current difficulty of controlling fusion reactions.