

AP Chemistry Correlation

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The chemical elements are fundamental building materials of matter, and all matter can be understood in terms of arrangements of atoms. These atoms retain their identity in chemical reactions.

Essential Knowledge	Pages	Illustrative Examples
1.A.1 Molecules are composed of specific combinations of atoms; different molecules are composed of combinations of different elements and of combinations of the same elements in differing amounts and proportions.	7–9, 39–40, 50–55, 76–83, 85–99, 121–126	Matter Map 9; Law of Multiple Proportions 40; Law of Conservation of Mass, 40; Monatomic Ions 51; Molecular Formulas 52; Mass Percent 85–88; Chemical Equations 94
1.A.2 Chemical analysis provides a method for determining the relative number of atoms in a substance, which can be used to identify the substance or determine its purity.	85–90, 149–151, 151–157	Empirical and Molecular Formulas 89–90; Gravimetric Analysis 149–151; Acid-Base Titration 151–154; Redox Titration 155–157
1.A.3 The mole is the fundamental unit for counting numbers of particles on the macroscopic level and allows quantitative connections to be drawn between laboratory experiments, which occur at the macroscopic level, and chemical processes, which occur at the atomic level.	76–104, 145–149	Average Atomic Mass 77; Mass/Mole/Number of Particles Conversions 79–83
1.B.1 The atom is composed of negatively charged electrons, which can leave the atom, and a positively charged nucleus that is made of protons and neutrons. The attraction of the electrons to the nucleus is the basis of the structure of the atom. Coulomb’s Law is qualitatively useful for understanding the structure of the atom.	40–46, 51, 282–287, 374–379	Cathode Ray Tube 4; Rutherford Expt 44; Photoelectric Effect 282–283
1.B.2 The electronic structure of the atom can be described using an electron configuration that reflects the concept of electrons in quantized energy levels or shells; the energetics of the electrons in the atom can be understood by consideration of Coulomb’s Law.	297–313	s, p, d Orbitals 297–300; Electron Configurations 309, 329, 333
1.C.1 Many properties of atoms exhibit periodic trends that are reflective of the periodicity of electronic structure.	48, 50, 310–314, 329–359, 369, 380–383, 957, 982–988, 995–998	Metals, Nonmetals, Metalloids 48; Electron Configuration Trends 313; Ground State Configurations 329, 333; Effective Nuclear Charge, Atomic and Ionic Radius 334–339; Ionization Energy 340–344; Electron Affinity 345–347; Lewis Dot Structure 369; Electronegativity 380
1.C.2 The currently accepted best model of the atom is based on the quantum mechanical model.	287, 295	
1.D.1 As is the case with all scientific models, any model of the atom is subject to refinement and change in response to new experimental results. In that sense, an atomic model is not regarded as an exact description of the atom, but rather a theoretical construct that fits a set of experimental data.	40–46, 277–281, 282–295, 334–347	Heisenberg’s Uncertainty Principle 292–293
1.D.2 An early model of the atom stated that all atoms of an element are identical. Mass spectrometry data demonstrate evidence that contradicts this early model.	83–85	Mass Spectrometer 83–85
1.D.3 The interaction of electromagnetic waves or light with matter is a powerful means to probe the structure of atoms and molecules, and to measure their concentration.	277–295	Properties of Waves 275–276; Electromagnetic Radiation 276–278; Photoelectric Effect 279–281; Spectroscopy 282–289; Ground vs Excited State 284; DeBroglie’s Equation 287–291
1.E.1 Physical and chemical processes can be depicted symbolically; when this is done, the illustration must conserve all atoms of all types.	90–95, 126–144, 178–184	

1.E.2 Conservation of atoms makes it possible to compute the masses of substances involved in physical and chemical processes. Chemical processes result in the formation of new substances, and the amount of these depends on the number and the types and masses of elements in the reactants, as well as the efficiency of the transformation.

90–104, 149–156,
193–195,
240–246, 730–739

Balancing Equations 94;
Gravimetric Analysis 150

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Chemical and physical properties of materials can be explained by the structure and the arrangements of atoms, ions, or molecules and the forces between them.

Essential Knowledge	Pages	Illustrative Examples
2.A.1 The different properties of solids and liquids can be explained by differences in their structures, both at the particulate level and in their supramolecular structures.	9, 10, 473–476, 486–491, 493–503	Properties of Gases, Liquids, and Solids 466; Surface Tension 473; Cohesion vs Adhesion 473; Viscosity 474
2.A.2 The gaseous state can be effectively modeled with a mathematical equation relating various macroscopic properties. A gas has neither a definite volume nor a definite shape; because the effects of attractive forces are minimal, we usually assume that the particles move independently.	7, 9, 173–213, 493–503	Pressure of Gases 174–177; Gas Laws 178–184; Ideal Gas Equation 184–193; Gas Stoichiometry 193–195; Dalton’s Law 196–201; Non-Ideal Gases 210–213
2.A.3 Solutions are homogenous mixtures in which the physical properties are dependent on the concentration of the solute and the strengths of all interactions among the particles of the solutes and solvent.	7–9, 119–121, 145–149, 519–526, 532	Molarity, Dilutions, Solution Prep 145–149; Types of Solutions 521; Fractional Distillation Apparatus 537
2.B.1 London dispersion forces are attractive forces present between all atoms and molecules. London dispersion forces are often the strongest net intermolecular force between large molecules.	468–470	
2.B.2 Dipole forces result from the attraction among the positive ends and negative ends of polar molecules. Hydrogen bonding is a strong type of dipole-dipole force.	119–121, 467–472	Ion Hydration 120; Intermolecular Forces 467–472
2.B.3 Intermolecular forces play a key role in determining the properties of substances, including biological structures and interactions.	173–174, 467–472	
2.C.1 In covalent bonding, electrons are shared between the nuclei of two atoms to form a molecule or polyatomic ion. Electronegativity differences between the two atoms account for the distribution of the shared electrons and the polarity of the bond.	52–54, 379–385, 425–431	Covalent Bonds and Compounds 377–380; Bond Length 379; Octet Rule 378; Ionic vs Covalent Compounds 382; Polar Covalent Bonds 382, 384–385
2.C.2 Ionic bonding results from the net attraction between oppositely charged ions, closely packed together in a crystal lattice.	54–55, 173, 370–377, 486–489	Ionic Bond 370; Lattice Energy 375
2.C.3 Metallic bonding describes an array of positively charged metal cores surrounded by a sea of mobile valence electrons.	491	
2.C.4 The localized electron bonding model describes and predicts molecular geometry using Lewis diagrams and the VSEPR model.	384–398, 413–443, 452–453	Drawing Lewis Structures 384–386; Formal Charges 387–389; Resonance 390; Exceptions to the Octet Rule 392–398; VSEPR 414; Orbital Overlap 432; Hybrid Orbitals 431–443; Multiple Bonds 441–443
2.D.1 Ionic solids have high melting points, are brittle, and conduct electricity only when molten or in solution.	119–121, 372–377, 379–380, 486–489	Lattice Energy 375; Ionic vs Covalent Compounds 379–380, 491
2.D.2 Metallic solids are good conductors of heat and electricity, have a wide range of melting points, and are shiny, malleable, ductile, and readily alloyed.	488–493, 939–941	Metallic Crystals 491
2.D.3 Covalent network solids generally have extremely high melting points, are hard, and are thermal insulators. Some conduct electricity.	490, 939–941	Covalent Crystals 490–491
2.D.4 Molecular solids with low molecular weight usually have low melting points and are not expected to conduct electricity as solids, in solution, or when molten.	379–382, 490	

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Changes in matter involve the rearrangement and/or reorganization of atoms and/or the transfer of electrons.

Essential Knowledge	Pages	Illustrative Examples
3.A.1 A chemical change may be represented by a molecular, ionic, or net ionic equation.	90–104, 119–144, 149–157	Molecular and Net Ionic Equations 123–126
3.A.2 Quantitative information can be derived from stoichiometric calculations that utilize the mole ratios from the balanced chemical equations. The role of stoichiometry in real-world applications is important to note, so that it does not seem to be simply an exercise done only by chemists.	90–104, 149–152, 155, 193–195, 730–739	Mole Method 97–99; Limiting Reactants 100–102; Reaction Yields 103–104; Gravimetric Analysis 149–151; Volumetric Analysis 153–157; Gas Stoichiometry 193–195
3.B.1 Synthesis reactions are those in which atoms and/or molecules combine to form a new compound. Decomposition is the reverse of synthesis, a process whereby molecules are decomposed, often by the use of heat.	132–144	Combination and Decomposition Reactions 137–138
3.B.2 In a neutralization reaction, protons are transferred from an acid to a base.	126–132, 155–156, 730–742	Brønsted-Lowry Acids and Bases 127–132
3.B.3 In oxidation-reduction (redox) reactions, there is a net transfer of electrons. The species that loses electrons is oxidized, and the species that gains electrons is reduced.	132–143, 155–156, 813–824, 832–848	Redox Reactions 132–144; 813–816; Redox Titrations 155–156
3.C.1 Production of heat or light, formation of a gas, and formation of a precipitate and/or a color change are possible evidences that a chemical change has occurred.	121–132, 149–151	Precipitation Reactions 121
3.C.2 Net changes in energy for a chemical reaction can be endothermic or exothermic.	232–234, 246–262, 400–404	Diagrams of Exothermic and Endothermic Reactions 233, 243
3.C.3 Electrochemistry shows the interconversion between chemical and electrical energy in galvanic and electrolytic cells.	818–827, 843–849	Galvanic Cells 817; Standard Reduction Potentials 821; Standard Cell Potentials 819, 822; Free Energy and Cell Potential 825; Electrolysis 843, 845, 847–848

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Rates of chemical reactions are determined by details of the molecular collisions.

Essential Knowledge	Pages	Illustrative Examples
4.A.1 The rate of a reaction is influenced by the concentration or pressure of reactants, the phase of the reactants and products, and environmental factors such as temperature and solvent.	563–587	Expressing Rates 563–568; Rates and Stoichiometry 568–570
4.A.2 The rate law shows how the rate depends on reactant concentrations.	571–587	Method of Initial Rates 575; Concentration vs. Time Studies 578, 581, 583; Half Life 586; Rate Law Summary 589
4.A.3 The magnitude and temperature dependence of the rate of reaction is contained quantitatively in the rate constant.	563–587, 870–874	Examples Involving Rate Constant 575–587
4.B.1 Elementary reactions can be unimolecular or involve collisions between two or more molecules.	588–593	Collision Frequency 590
4.B.2 Not all collisions are successful. To get over the activation energy barrier, the colliding species need sufficient energy. Also, the orientations of the reactant molecules during the collision must allow for the rearrangement of reactant bonds to form product bonds.	588–589	Activation Energy 590, Orientation 593
4.B.3 A successful collision can be viewed as following a reaction path with an associated energy profile.	588–593	Reaction Pathway 600

4.C.1 The mechanism of a multistep reaction consists of a series of elementary reactions that add up to the overall reaction.	589–607	Reaction Mechanisms 594–598
4.C.2 In many reactions, the rate is set by the slowest elementary reaction, or rate-limiting step.	594–607	Rate-determining Step 594–598
4.C.3 Reaction intermediates, which are formed during the reaction but not present in the overall reaction, play an important role in multistep reactions.	594–595	Intermediates 596–597
4.D.1 Catalysts function by lowering the activation energy of an elementary step in a reaction mechanism, and by providing a new and faster reaction mechanism.	599–607	Activation Energy and Catalysts 600
4.D.2 Important classes in catalysis include acid-base catalysis, surface catalysis, and enzyme catalysis.	599–607	Surface and Enzyme Catalysis 601–607

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The laws of thermodynamics describe the essential role of energy and explain and predict the direction of changes in matter.

Essential Knowledge	Pages	Illustrative Examples
5.A.1 Temperature is a measure of the average kinetic energy of atoms and molecules.	202–207, 493–503	Maxwell-Boltzmann Distribution Curve 205
5.A.2 The process of kinetic energy transfer at the particulate scale is referred to in this course as heat transfer, and the spontaneous direction of the transfer is always from a hot to a cold body.	234–240, 246–254	
5.B.1 Energy is transferred between systems either through heat transfer or through one system doing work on the other system.	232–254	Work and Heat 236–239
5.B.2 When two systems are in contact with each other and are otherwise isolated, the energy that comes out of one system is equal to the energy that goes into the other system. The combined energy of the two systems remains fixed. Energy transfer can occur through either heat exchange or work.	231–254, 777	Work and Heat 236–239
5.B.3 Chemical systems undergo three main processes that change their energy: heating/cooling, phase transitions, and chemical reactions.	234–252, 258–261, 493–503	Heating Curve 501; Energy of Warming/Phase Changes 502
5.B.4 Calorimetry is an experimental technique that is used to measure the change in energy of a chemical system.	246–252	Use of Calorimeter 246–252
5.C.1 Potential energy is associated with a particular geometric arrangement of atoms or ions and the electrostatic interactions between them.	231, 398–402	Bond Energies 399
5.C.2 The net energy change during a reaction is the sum of the energy required to break the bonds in the reactant molecules and the energy released in forming the bonds of the product molecules. The net change in energy may be positive for endothermic reactions where energy is required, or negative for exothermic reactions where energy is released.	232–234, 253–261, 400–402	Enthalpy of Reactions Using Bond Energies 400–402; Hess' Law 258–260
5.D.1 Potential energy is associated with the interaction of molecules; as molecules draw near each other, they experience an attractive force.	231, 467–473, 493–503	Intermolecular Attractions 467–473
5.D.2 At the particulate scale, chemical processes can be distinguished from physical processes because chemical bonds can be distinguished from intermolecular interactions.	232–234, 243, 258–261, 493–503	
5.D.3 Noncovalent and intermolecular interactions play important roles in many biological and polymer systems.	1065–1077	Hydrogen Bonding: Proteins 1071–1073, DNA 1077
5.E.1 Entropy is a measure of the dispersal of matter and energy.	778–787	Entropy 783 Entropy of Reactions 784–786
5.E.2 Some physical or chemical processes involve both a decrease in the internal energy of the components ($\Delta H^\circ < 0$) under consideration <i>and</i> an increase in the entropy of those components ($\Delta S^\circ > 0$). These processes are necessarily “thermodynamically favored” ($\Delta G^\circ < 0$).	777–778, 783–787, 789–796	“Thermodynamically Favored” Process 777–778, 790

5.E.3 If a chemical or physical process is not driven by <i>both</i> entropy and enthalpy changes, then the Gibbs free energy change can be used to determine whether the process is thermodynamically favored.	789–796	Gibbs Free Energy Change 792–793
5.E.4 External sources of energy can be used to drive change in cases where the Gibbs free energy change is positive.	789–802, 841–848	Coupling to Drive a Reaction 800–802
5.E.5 A thermodynamically favored process may not occur due to kinetic constraints (kinetic vs. thermodynamic control).	777–778, 789–796	

BIG IDEA 6 Any bond or intermolecular attraction that can be formed can be broken. These two processes are in a dynamic competition, sensitive to initial conditions and external perturbations.

Essential Knowledge	Pages	Illustrative Examples
6.A.1 In many classes of reactions, it is important to consider both the forward and reverse reaction.	119–121, 126–132, 493–503, 622–636	Reversible Reactions 624–625
6.A.2 The current state of a system undergoing a reversible reaction can be characterized by the extent to which reactants have been converted to products. The relative quantities of reaction components are quantitatively described by the reaction quotient, Q .	622–636, 638–654	Finding and Using Equilibrium Constants (Keq) 627–630, 633–634, 641–644; Manipulating Keq 634–636, Reaction Quotient (Q) and Keq 638–640, 648–650
6.A.3 When a system is at equilibrium, all macroscopic variables, such as concentrations, partial pressures, and temperature, do not change over time. Equilibrium results from an equality between the rates of the forward and reverse reactions, at which point $Q = K$.	622–636, 638–654	See Illustrative Examples for 6.A.2
6.A.4 The magnitude of the equilibrium constant, K , can be used to determine whether the equilibrium lies toward the reactant side or product side.	624–638, 640–646, 689–694	Keq vs. Reactants and Products 627, 676
6.B.1 Systems at equilibrium respond to disturbances by partially countering the effect of the disturbance (LeChatelier’s principle).	646–656	Predicting Shifts: Concentration 645–646, Pressure 646–648, Temperature and Catalyst 648–650
6.B.2 A disturbance to a system at equilibrium causes Q to differ from K , thereby taking the system out of the original equilibrium state. The system responds by bringing Q back into agreement with K , thereby establishing a new equilibrium state.	646–654, 721–724	Keq, Q and LeChatelier 646–654
6.C.1 Chemical equilibrium reasoning can be used to describe the proton-transfer reactions of acid-base chemistry.	669–703, 741–743	Strong vs. Weak Acid 674; Kw 669; pH for Strong and Weak Acids/Bases 676, 613–683, 685–687; pH of Polyprotic Acids 691–693; pH of Salts 696–701, pH Titrations 730–739
6.C.2 The pH is an important characteristic of aqueous solutions that can be controlled with buffers. Comparing pH to pK_a allows one to determine the protonation state of a molecule with a labile proton.	670–688, 724–730	Buffer Recognition 726, pH in Buffer Solutions 722, 726–728, Buffer Prep 729–730
6.C.3 The solubility of a substance can be understood in terms of chemical equilibrium.	721, 742–763	Solubility Product Problems 745–749
6.D.1 When the difference in Gibbs free energy between reactants and products (ΔG°) is much larger than the thermal energy (RT), the equilibrium constant is either very small (for $\Delta G^\circ < 0$) or very large (for $\Delta G^\circ > 0$). When ΔG° is comparable to the thermal energy (RT), the equilibrium constant is near 1.	796–800	Free Energy and Keq 798–800