AP Chemistry

Name ____________________________

Preparation for AP Chemistry

Period _______ Date ______________

S U M M E R   A S S I G N M E N T

☐ **Get a copy of AP Chemistry Princeton Review.** 2016 Edition will be fine.

☐ **Familiarize yourself with the Periodic Table.** You should recognize quickly the symbols (name them) and their positions on the Periodic Table of the first 20 elements. The last page of this packet has a Periodic Table for your reference.

☐ **Memorize the list of sixty familiar ions.** You should be able to write the symbol of the ion given the name, and vice versa.

- Included in this packet are: (1) a list of ions by name, (2) a list of ions by grouping, (2) and (3) four practice ions quizzes.
- You should be able to complete the ions quiz in less than four minutes without trouble. (Basically you need to know your ions better than you know your multiplication table.)
- There will be a quiz in the first week of school.
- You may want to print out and cut the flashcards of the ions, and then arrange them by groups based on patterns you observe. (For example, which ones are made of only one element / several elements? Which ones have positive charges vs. negative charges? If it's an ion from a single element, where are they on the periodic table? Which ones have several possible charges? How are they named, e.g. –ium, -ide, -ate, -ite, hypo-, per-, bi-, thio-, -ous, -ic?)

☐ **Learn every topic.** Refer to the list of Review Topics. These are all the concepts and skills that you are expected to have learned during your first year of chemistry and that you can do with ease. All of these will be covered during the summer Pre-AP Chemistry class.

- Review the relevant sections in Princeton Review

☐ **Complete the Review Problems. THIS IS NOT HOMEWORK** Please complete these problems neatly, showing your work when appropriate. All of these are the types of problems with which you should be very familiar. This assignment will take a while, so don't put it off! The better you are able to do these problems, the more prepared you will be to succeed in AP Chemistry.

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AP Chemistry

Name ______________________________

Preparation for AP Chemistry

Period ________ Date ________________

IONS LIST BY NAME

Acetate	$C_2H_3O_2^-$ or CH_3COO^-	Hydrogen sulfate, bisulfate	HSO_4^-	Oxide	O^{2-}
Aluminum	Al^{3+}	Hydronium	H_3O^+	Perbromate	BrO_4^-
Ammonium	NH_4^+	Hydroxide	OH^-	Perchlorate	ClO_4^-
Barium	Ba^{2+}	Hypobromite	BrO^- or OBr^-	Periodate	IO_4^-
Bromate	BrO_3^-	Hypochlorite	ClO^- or OCl^-	Permanganate	MnO_4^-
Bromide	Br^-	Hypoiodite	IO^- or OI^-	Peroxide	O_2^{2-}
Bromite	BrO_2^-	Iodate	IO_3^-	Phosphate	PO_4^{3-}
Calcium	Ca^{2+}	Iodide	I^-	Phosphite	PO_3^{3-}
Carbonate	CO_3^{2-}	Iodite	IO_2^-	Potassium	K^+
Chlorate	ClO_3^-	Iron (II), ferrous	Fe^{2+}	Silver	Ag^+
Chloride	Cl^-	Iron (III), ferric	Fe^{3+}	Sodium	Na^+
Chlorite	ClO_2^-	Lead (II), plumbous	Pb^{2+}	Strontium	Sr^{2+}
Chromate	CrO_4^{2-}	Lead (IV), plumbic	Pb^{4+}	Sulfate	SO_4^{2-}
Copper (I), cuprous	Cu^+	Lithium	Li^+	Sulfide	S^{2-}
Copper (II), cupric	Cu^{2+}	Magnesium	Mg^{2+}	Sulfite	SO_3^{2-}
Cyanide	CN^-	Mercury (I), mercurous	Hg_2^{2+}	Thiocyanate	SCN^-
Dichromate	$Cr_2O_7^{2-}$	Mercury (II), mercuric	Hg^{2+}	Thiosulfate	$S_2O_3^{2-}$
Fluoride	F^-	Nickel	Ni^{2+}	Tin (II), stannous	Sn^{2+}
Hydrogen	H^+	Nitrate	NO_3^-	Tin (IV), stannic	Sn^{4+}
Hydrogen carbonate, bicarbonate,	HCO_3^-	Nitrite	NO_2^-	Zinc	Zn^{2+}

AP Chemistry

Name ________________________________

Preparation for AP Chemistry

Period _________ Date _________________

IONS LIST BY SET

Set 1

Aluminum	Al^{3+}	Fluoride	F^{-}	Oxide	O^{2-}
Barium	Ba^{2+}	Hydrogen	H^{+}	Potassium	K^{+}
Bromide	Br^{-}	Iodide	I^{-}	Sodium	Na^{+}
Calcium	Ca^{2+}	Lithium	Li^{+}	Strontium	Sr^{2+}
Chloride	Cl^{-}	Magnesium	Mg^{2+}	Sulfide	S^{2-}

Set 2

Ammonium	NH_4^{+}	Iron (III), ferric	Fe^{3+}	Nickel	Ni^{2+}
Copper (I), cuprous	Cu^{+}	Lead (II), plumbous	Pb^{2+}	Silver	Ag^{+}
Copper (II), cupric	Cu^{2+}	Lead (IV), plumbic	Pb^{4+}	Tin (II), stannous	Sn^{2+}
Hydronium	H_3O^{+}	Mercury (I), mercurous	Hg_2^{2+}	Tin (IV), stannic	Sn^{4+}
Iron (II), ferrous	Fe^{2+}	Mercury (II), mercuric	Hg^{2+}	Zinc	Zn^{2+}

Set 3

Acetate	$C_2H_3O_2^{-}$ or CH_3COO^{-}	Chlorite	ClO_2^{-}	Nitrite	NO_2^{-}
Bromate	BrO_3^{-}	Hydroxide	OH^{-}	Phosphate	PO_4^{3-}
Bromite	BrO_2^{-}	Iodate	IO_3^{-}	Phosphite	PO_3^{3-}
Carbonate	CO_3^{2-}	Iodite	IO_2^{-}	Sulfate	SO_4^{2-}
Chlorate	ClO_3^{-}	Nitrate	NO_3^{-}	Sulfite	SO_3^{2-}

Set 4

Chromate	CrO_4^{2-}	Hypobromite	BrO^{-} or OBr^{-}	Periodate	IO_4^{-}
Cyanide	CN^{-}	Hypochlorite	ClO^{-} or OCl^{-}	Permanganate	MnO_4^{-}
Dichromate	$Cr_2O_7^{2-}$	Hypoiodite	IO^{-} or OI^{-}	Peroxide	O_2^{2-}
Hydrogen carbonate, bicarbonate,	HCO_3^{-}	Perbromate	BrO_4^{-}	Thiocyanate	SCN^{-}
Hydrogen sulfate, bisulfate	HSO_4^{-}	Perchlorate	ClO_4^{-}	Thiosulfate	$S_2O_3^{2-}$

IONS SET 1 **Aluminum**	IONS SET 1 **Barium**	IONS SET 1 **Bromide**	IONS SET 1 **Calcium**	IONS SET 1 **Chloride**
IONS SET 1 **Fluoride**	IONS SET 1 **Hydrogen**	IONS SET 1 **Iodide**	IONS SET 1 **Lithium**	IONS SET 1 **Magnesium**
IONS SET 1 **Oxide**	IONS SET 1 **Potassium**	IONS SET 1 **Sodium**	IONS SET 1 **Strontium**	IONS SET 1 **Sulfide**
IONS SET 2 **Ammonium**	IONS SET 2 **Copper (I), cuprous**	IONS SET 2 **Copper (II), cupric**	IONS SET 2 **Hydronium**	IONS SET 2 **Iron (II), ferrous**

Cl^{-}	Ca^{2+}	Br^{-}	Ba^{2+}	Al^{3+}
Mg^{2+}	Li^{+}	I^{-}	H^{+}	F^{-}
S^{2-}	Sr^{2+}	Na^{+}	K^{+}	O^{2-}
Fe^{2+}	H_3O^{+}	Cu^{2+}	Cu^{+}	NH_4^{+}

IONS SET 2 **Iron (III), ferric**	IONS SET 2 **Lead (II), plumbous**	IONS SET 2 **Lead (IV), plumbic**	IONS SET 2 **Mercury (I), mercurous**	IONS SET 2 **Mercury (II), mercuric**
IONS SET 2 **Nickel**	IONS SET 2 **Silver**	IONS SET 2 **Tin (II), stannous**	IONS SET 2 **Tin (IV), stannic**	IONS SET 2 **Zinc**
IONS SET 3 **Acetate**	IONS SET 3 **Bromate**	IONS SET 3 **Bromite**	IONS SET 3 **Carbonate**	IONS SET 3 **Chlorate**
IONS SET 3 **Chlorite**	IONS SET 3 **Hydroxide**	IONS SET 3 **Iodate**	IONS SET 3 **Iodite**	IONS SET 3 **Nitrate**

Hg^{2+}	Hg_2^{2+}	Pb^{4+}	Pb^{2+}	Fe^{3+}
Zn^{2+}	Sn^{4+}	Sn^{2+}	Ag^{+}	Ni^{2+}
ClO_3^-	CO_3^{2-}	BrO_2^-	BrO_3^-	$C_2H_3O_2^-$ or CH_3COO^-
NO_3^-	IO_2^-	IO_3^-	OH^-	ClO_2^-

IONS SET 3 **Nitrite**	IONS SET 3 **Phosphate**	IONS SET 3 **Phosphite**	IONS SET 3 **Sulfate**	IONS SET 3 **Sulfite**
IONS SET 4 **Chromate**	IONS SET 4 **Cyanide**	IONS SET 4 **Dichromate**	IONS SET 4 **Hydrogen carbonate, bicarbonate**	IONS SET 4 **Hydrogen sulfate, bisulfate**
IONS SET 4 **Hypobromite**	IONS SET 4 **Hypochlorite**	IONS SET 4 **Hypoiodite**	IONS SET 4 **Perbromate**	IONS SET 4 **Perchlorate**
IONS SET 4 **Periodate**	IONS SET 4 **Permanganate**	IONS SET 4 **Peroxide**	IONS SET 4 **Thiocyanate**	IONS SET 4 **Thiosulfate**

SO_3^{2-}	SO_4^{2-}	PO_3^{3-}	PO_4^{3-}	NO_2^-
HSO_4^-	HCO_3^-	$Cr_2O_7^{2-}$	CN^-	CrO_4^{2-}
ClO_4^-	BrO_4^-	IO^- or OI^-	ClO^- or OCl^-	BrO^- or OBr^-
$S_2O_3^{2-}$	SCN^-	O_2^{2-}	MnO_4^-	IO_4^-

AP Chemistry

Name ______________________________

Preparation for AP Chemistry

Period ________ Date ________________

PRACTICE IONS QUIZ - 1

Name	Formula
Hypobromite	
Iodate	
Permanganate	
Sodium	
Ammonium	
Bisulfate	
Phosphite	
Iodite	
Silver	
Cuprous	
Phosphate	
Iodide	
Sulfite	
Lithium	
Calcium	
Sulfide	
Acetate	
Mercurous	
Bromite	
Aluminum	

Name	Formula
Hydrogen	
Thiocyanate	
Perbromate	
Nitrite	
Hypoiodite	
Nickel	
Chlorate	
Stannic	
Oxide	
Fluoride	
Cyanide	
Thiosulfate	
Hydroxide	
Mercuric	
Cupric	
Hypochlorite	
Plumbic	
Stannous	
Plumbous	
Carbonate	

Name	Formula
Dichromate	
Nitrate	
Bromate	
Bicarbonate,	
Chromate	
Hydronium	
Peroxide	
Ferric	
Chloride	
Strontium	
Periodate	
Potassium	
Chlorite	
Ferrous	
Bromide	
Perchlorate	
Sulfate	
Zinc	
Barium	
Magnesium	

AP Chemistry

Name ________________________________

Preparation for AP Chemistry

Period _________ Date _________________

PRACTICE IONS QUIZ - 2

Name	Formula
Ferrous	
Plumbous	
Calcium	
Bromite	
Thiosulfate	
Sodium	
Chlorite	
Chromate	
Strontium	
Sulfite	
Stannous	
Hydroxide	
Nitrate	
Cyanide	
Aluminum	
Sulfide	
Bromide	
Magnesium	
Hypobromite	
Iodide	

Name	Formula
Hydronium	
Nitrite	
Iodite	
Perbromate	
Bicarbonate,	
Zinc	
Barium	
Perchlorate	
Peroxide	
Periodate	
Mercuric	
Silver	
Phosphate	
Thiocyanate	
Bromate	
Carbonate	
Bisulfate	
Chloride	
Hydrogen	
Cuprous	

Name	Formula
Iodate	
Cupric	
Phosphite	
Oxide	
Chlorate	
Hypochlorite	
Permanganate	
Dichromate	
Plumbic	
Nickel	
Mercurous	
Fluoride	
Potassium	
Sulfate	
Lithium	
Ammonium	
Acetate	
Stannic	
Hypoiodite	
Ferric	

AP Chemistry

Name ______________________________

Preparation for AP Chemistry

Period ________ Date ________________

PRACTICE IONS QUIZ - 3

Name	Formula
Perchlorate	
Iodide	
Plumbous	
Calcium	
Thiocyanate	
Bromate	
Nitrite	
Sulfide	
Chromate	
Sulfite	
Sulfate	
Phosphate	
Phosphite	
Barium	
Sodium	
Hypoiodite	
Cupric	
Periodate	
Aluminum	
Stannous	

Name	Formula
Cuprous	
Zinc	
Perbromate	
Bicarbonate,	
Oxide	
Stannic	
Nickel	
Hydronium	
Ferrous	
Hypobromite	
Iodite	
Lithium	
Ammonium	
Strontium	
Dichromate	
Plumbic	
Acetate	
Magnesium	
Iodate	
Bromite	

Name	Formula
Mercurous	
Silver	
Thiosulfate	
Fluoride	
Ferric	
Hypochlorite	
Potassium	
Nitrate	
Peroxide	
Cyanide	
Mercuric	
Hydrogen	
Permanganate	
Chlorate	
Chloride	
Chlorite	
Carbonate	
Hydroxide	
Bisulfate	
Bromide	

AP Chemistry

Name ______________________________

Preparation for AP Chemistry

Period ________ Date ________________

PRACTICE IONS QUIZ - 4

Name	Formula
Aluminum	
Perbromate	
Sulfide	
Cuprous	
Nitrite	
Calcium	
Potassium	
Sulfite	
Hydroxide	
Chlorite	
Chlorate	
Zinc	
Mercuric	
Stannous	
Periodate	
Bromide	
Magnesium	
Cyanide	
Barium	
Ammonium	

Name	Formula
Sodium	
Bromite	
Lithium	
Chloride	
Thiocyanate	
Nickel	
Oxide	
Bromate	
Sulfate	
Hydronium	
Bicarbonate,	
Hydrogen	
Thiosulfate	
Silver	
Perchlorate	
Mercurous	
Iodite	
Iodate	
Hypobromite	
Plumbic	

Name	Formula
Peroxide	
Phosphate	
Hypochlorite	
Nitrate	
Phosphite	
Ferric	
Strontium	
Fluoride	
Stannic	
Chromate	
Ferrous	
Plumbous	
Carbonate	
Bisulfate	
Hypoiodite	
Permanganate	
Iodide	
Cupric	
Dichromate	
Acetate	

AP Chemistry

Name ____________________________

Preparation for AP Chemistry

Period _______ Date ______________

R E V I E W   T O P I C S

Big Picture Topics

Topic 1 – Foundations of Chemistry

0.1 Experimental Design
- Lab safety
- Stating researchable questions, observations, claims and evidence

1.1 Measurements
- Uncertainty and significant figures
- Scientific notation
- Precision vs. accuracy

1.2 Dimensional Analysis
- Metric System
- Temperature (Celsius and kelvin)
- Dimensional analysis

Topic 2 – Elements and Compounds

2.1 Atoms, Ions, Isotopes
- Historical models of atoms
- Isotopic notation of atom
- Atomic mass

2.2 Classification of Matter
- States of matter
- Classifying substances
- Periodic Table
- Names of common elements and ions

2.3 Compounds
- Types of Compounds
- Equations of formation and decomposition
- Writing and naming ionic compounds, covalent compounds, and acids

2.4 Moles
- Molar mass
- Moles of particles ↔ # particles ↔ mass of sample ↔ volume of gas @ STP
- Empirical and molecular formula
- Chemical analysis

Topic 3 – Organic Chemistry

3.1 Hydrocarbons
- Alkanes, alkenes, alkynes
- Isomers

3.2 Functional Groups

3.3 Organic Reactions
- Condensation (Esterification)
- Polymerization
- Combustion

Topic 4 – Salt and Solutions

4.1 Solubility
- Saturated vs. unsaturated solutions
- Soluble vs. insoluble salts
- Strong vs. weak acids/bases
- Factors affecting solubility (temperature, pressure)

4.2 Concentration
- Molarity
- Dilutions

4.3 Properties of Solutions
- Electrolytic properties
- Acid/base properties
- Arrhenius and Bronsted-Lowry ideas of acids/bases
- pH and pOH

4.4 Describing Solutions
- Find concentrations of strong acid, strong base, soluble salt solutions.
- Find concentrations and pH of weak acid and weak base solutions. (K_a and K_b)
- Find concentrations of insoluble salt solutions. (K_{sp})

Topic 5 – Chemical Reactions

5.1 Types of Reactions
- Classifying reactions
- Balancing equations
- Predicting products of Synthesis and Decomposition reactions

5.2 Reactions in Aqueous Solutions
- Net ionic equations
- Double replacement reactions

5.3 Redox Reactions
- Oxidation and reduction ideas
- Single replacement reactions

5.4 Stoichiometry

5.5 Limiting Reactants
- Limiting and excess reactant stoichiometry
- Percent yield

Topic 6 – Thermochemistry

6.1 Heat
- Endothermic and exothermic processes
- Calorimetry

6.2 Changes in State
- Enthalpies of fusion and vaporization
- Heating curve

6.3 Enthalpy of Reactions
- Enthalpy vs. Heat
- Products – Reactants
- Bond energies

Topic 7 – Kinetics

7.1 Collision Theory
- Collision Theory
- Potential energy graph and kinetic energy distribution
- Catalysts
- Multi-step-reactions and reaction mechanisms

7.2 Rates
- Concentration vs. Time graph
- Reaction rate
- Rate Law using initial rate data

Topic 8 – Equilibrium

8.1 Equilibrium Ideas
- Reversible vs. Irreversible Reactions
- Equilibrium condition
- Equilibrium Constant, K_{eq}
- Reaction Quotient, Q

8.2 Equilibrium Problems
- ICE Box Problems

8.3 Le Chatelier's Principle

Topic 9 – Atomic Structure

9.1 Atomic Orbitals
- Quantized Energy Levels
- Atomic orbitals: s, p, d, f

9.2 Electron Configuration
- Long form, short form
- Electron Configuration of atoms and ions
- Valence electrons and ions formed

9.3 Periodic Trends
- Atomic size
- Ionization Energy and Electronegativity
- Successive ionization energies
- Atomic vs. Ionic radius
- Radius of isoelectronic species

Topic 10 – Molecular Structure

10.1 Types of Bonds
- Types of Bonds (Non-polar covalent, polar covalent, ionic, metallic)
- Lewis Structures of Atoms and Monoatomic ions
- Coulomb's Law and Strength of Ionic Bonds

10.2 Lewis Structures
- Lewis Structures – simple covalent compounds, multiple bonds, expanded octets, resonance
- Bond order, bond energy, bond length

10.3 Molecular Shapes
- Hybridization
- Shapes of compounds with steric numbers 2, 3, and 4
- Polarity of Compounds

10.4 Inter-Molecular Forces
- IMFs
- Types of Solids

Topic 11 – Properties of Gases

11.1 Kinetic Molecular Theory
- Kinetic Molecular Theory
- Units of Pressure
- Graham's Law of Effusion

11.2 Gas Laws
- Boyle', Charles', Gay-Lussac's, Avogadro's Laws
- Ideal Gas Law

11.3 Gas Mixtures
- Dalton's Law of Partial Pressures

Detailed Objectives

Topic		Objectives	I Can
Topic 1 – Chemistry Foundations			
0.1 Experimental Design	0.1.1	Read an MSDS to identify physical properties, chemical properties, and safety precautions for a chemical.	
	0.1.2	State potential hazards and safety procedures when performing an experiment in a lab.	
	0.1.3	Identify the controlled variables, dependent variable, and independent variable in an experiment.	
	0.1.4	Write a researchable question about relating the dependent and independent variables.	
	0.1.5	Make detailed qualitative observations when performing an experiment.	
	0.1.6	Create a graph that relates the dependent and independent variables.	
	0.1.7	Make claims and justify them with evidence from the observations or data.	
1.1 Measurements	1.1.1	Appropriately read and report a measurement correctly with one uncertain digit.	
	1.1.2	State the number of significant digits in a measurement.	
	1.1.3	Identify a number as a measurement or an exact (or defined) value, and state that exact values have infinite significant digits.	
	1.1.4	Write a number in both scientific notation and standard decimal notation.	
	1.1.5	Carry out arithmetic operations (i.e. multiplication, division, addition, subtraction) with numbers in standard decimal notation and scientific notation, reporting the answer with the correct number of significant figures.	
	1.1.6	Report a set of measurements or calculations by writing: {average value} ± {average deviation}	
	1.1.7	Find the percent error a measurement.	
	1.1.8	Describe the accuracy or precision of a measurement or calculation.	
1.2 Dimensional Analysis	1.2.1	Identify the appropriate Topics for a measurement (e.g. volume, mass, length, time, temperature).	
	1.2.2	Write a derived Topic in terms of base metric Topics.	
	1.2.3	Convert measurements between metric quantities (e.g. 30 mg = ? dg).	
	1.2.4	Convert volumes between L and m^3.	
	1.2.5	Convert between kelvin and Celsius temperature scales.	
	1.2.6	Compare the relative average kinetic energy between two samples from the kelvin temperature.	
	1.2.7	Use dimensional analysis to convert like Topics (e.g. length → length)	
	1.2.8	Use dimensional analysis with unlike relationships (ratios and rates) to convert unlike Topics.	
1.3 Physical Properties	1.3.1	State physical and chemical properties of a substance.	
	1.3.2	Represent an observed change using a chemical equation.	
	1.3.3	Determine whether a change (or process) is considered physical or chemical.	
	1.3.4	Show work for a calculation using a formula using I.E.S.A.	
	1.3.5	Compare the mass, volume or density of two objects qualitatively and conceptually.	
	1.3.6	Calculate the mass, volume or density of an object using the definition of density.	
Topic 2 – Elements and Compounds			
2.1 Atoms, Ions, Isotopes	2.1.1	Describe the models of the atom proposed by Dalton, Thomson, and Rutherford, and explain the limitations and/or inaccuracies of each.	2.1, 2.2
	2.1.2	Describe the composition of an atom and the features (relative mass, size, charge, location) of subatomic particles (protons, neutrons, and electrons).	2.3, 2.4
	2.1.3	Describe an atom using isotopic notation. Find the number of protons, neutrons, and electrons of an atom using atomic number, mass number, and charge. Explain the relationship between two atoms with different protons, neutrons, or electrons.	

Topic	Objectives		
	2.1.4	Write the average atomic mass of an element and explain that it represents the average mass of naturally occurring isotopes.	
	2.1.5	Calculate the average atomic mass of an element using abundance data of isotopes.	
2.2 Classification of Matter	2.2.1	Describe the three states of matter (solid, liquid, and gas) macroscopically (i.e. their shape and volume) and molecularly (in terms of interparticle attractions and kinetic molecular theory).	
	2.2.2	Categorize matter into mixtures and pure substances, and pure substances into elements and compounds.	
	2.2.3	Describe how the periodic table is organized in periods and families/groups.	
	2.2.4	Categorize an element as a metal, metalloid, or non-metal.	
	2.2.5	Describe typical properties of metals and non-metals.	
	2.2.6	Categorize elements in families (alkali/ alkaline earth/ transition/ inner transition/ halogens/ noble gases).	
	2.2.7	Describe the similarities and differences between elements in the same family (e.g. similar chemical properties, monatomic ions formed) and same period (e.g. similar size).	
	2.2.8	Recall the name and symbol of 40 common elements and 60 monatomic and polyatomic ions.	
	2.2.9	Identify the states of elements in standard state: gases (noble gases + N_2, O_2, F_2, Cl_2), liquids (Br_2 and Hg), solids (the rest).	
	2.2.10	list the seven diatomic elements (H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2).	
2.3 Compounds	2.3.1	Categorize a compound as an ionic compound, covalent compound, organic compound, or acid.	
	2.3.2	Given the formula of a compound, determine the number of each element in the compound.	
	2.3.3	Write the balanced equation for the formation of a compound (synthesis) or the decomposition of a compound.	
	2.3.4	Write the molecular formula of an ionic compound given a cation and anion.	
	2.3.5	Decompose an ionic compound into its constituent cation and anion.	
	2.3.6	Write the name and formula of an ionic compound.	
	2.3.7	Write the name and formula of a binary covalent compound.	
	2.3.8	Write the name and formula of an acid.	
2.4 Moles	2.4.1	Calculate the molecular mass of a compound and the percent composition of each element in a compound.	
	2.4.2	Describe a mole as 6.022×10^{23} objects.	
	2.4.3	Calculate the molar mass of a compound given the mass and number of moles of the compound in a sample.	
	2.4.4	Find the mass (using the molar mass), number of particles (using 1 mole = 6.022×10^{23} objects), number of moles, and volume of a gas at STP (using 1 mole = 22.4 L) of a sample using dimensional analysis.	
	2.4.5	Determine the empirical formula of a compound given percent composition of each element.	
	2.4.6	Determine the molecular formula of a compound given the empirical formula and molar mass.	
	2.4.7	Use data from the synthesis, decomposition, or combustion data to determine the empirical formula of a compound.	
Topic 3 – Organic Chemistry			

<table>
<tr><th>Topic</th><th colspan="2">Objectives</th><th>I Can</th></tr>
<tr><td rowspan="3">3.1 Hydrocarbons</td><td>3.1.1</td><td>Represent a hydrocarbon (alkane, alkene, or alkyne) using its molecular formula and structural formula.</td><td rowspan="1"></td></tr>
<tr><td>3.1.2

3.1.3
3.1.4</td><td>Describe characteristics of alkanes, alkenes, and alkynes. To do this, I can:
• State that
<table><tr><td>alkanes</td><td>all single bonds</td><td>end with “-ane”</td><td>follow the formula C_nX_{2n+2}</td></tr><tr><td>alkenes</td><td>one double bond</td><td>end with “-ene”</td><td>follow the formula C_nX_{2n}</td></tr><tr><td>alkynes</td><td>one triple bond</td><td>end with “-yne”</td><td>follow the formula C_nX_{2n-2}</td></tr></table>(X represents H, F, Cl, Br, or I)
• Know the bond angle of H−C−C bond for alkanes, alkenes, and alkynes.
• Demonstrate that double and triple bonds cannot rotate like a single bond.
• State that “saturated” means “saturated with hydrogens” and describes alkanes. State that alkenes, alkynes, and cyclic hydrocarbons are all “unsaturated.”
Given a formula, recognize whether the molecule is an alkane, alkene, or alkyne.
Write the name of a hydrocarbon given its structural formula, and vice versa.</td><td></td></tr>
<tr><td>3.1.5
3.1.6
3.1.7

3.1.8
3.1.9</td><td>Recognize whether two compounds are isomers.
Draw and name isomers of unsubstituted alkanes by rearranging the carbon skeleton.
Draw and name isomers of substituted hydrocarbons, by moving the side group to a different carbon.
Draw and name isomers of alkenes and alkynes by moving the double or triple bond.
Demonstrate cis- and trans- isomerism using dichloroeth<u>ene</u>, $C_2H_2Cl_2$.</td><td></td></tr>
<tr><td>3.2 Functional Groups</td><td>3.2.1

3.2.2

3.2.3</td><td>Identify the functional group in a compound. (i.e. alcohol, ether, aldehyde, ketone, carboxylic acid ester, amine)
Write the structural formula give the name of a compound that contains an oxygen- or nitrogen-containing functional group.
State examples of these compounds in daily life.</td><td></td></tr>
<tr><td rowspan="3">3.3 Organic Reactions</td><td>3.3.1
3.3.2

3.3.3</td><td>Show how compounds can combine by removing water: a “condensation reaction.”
State that ester formation (esterification) consists of:
Carboxylic Acid + Alcohol → Ester + H_2O
Given a carboxylic acid and alcohol, write the products of an esterification reaction.
State that protein synthesis involves:
Amino Acid + Amino Acid → Protein + H_2O</td><td></td></tr>
<tr><td>3.3.4

3.3.5</td><td>Know that a polymer (“poly” means many, “mer” means parts) consists of many repeating parts (monomers).
State natural examples of polymers such as proteins (many amino acids), carbohydrates (many sugar Topics), and Nylon (many difunctional monomers).</td><td></td></tr>
<tr><td>3.3.6
3.3.7

3.3.8</td><td>State that combustion is another name for burning.
State the basic equation for the combustion reaction:
fuel + O_2 → CO_2 + H_2O
Write the balanced equation for the combustion of a given organic compound.</td><td></td></tr>
<tr><th colspan="4">Topic 4 – Salts and Solutions</th></tr>
<tr><td rowspan="3">4.1 Solubility</td><td>4.1.1
4.1.2
4.1.3
4.1.4</td><td>Describe a solution as a homogeneous mixture of a solute and a solvent.
Determine whether a solution is saturated.
Predict whether two substances are miscible using the "like dissolves like" idea.
Determine whether a substance is a salt, an acid, or a base in an aqueous solution.</td><td></td></tr>
<tr><td>4.1.5
4.1.6
4.1.7</td><td>Describe how ionic compounds dissociate into ions when they dissolve in solution.
Use the solubility rules to determine whether a salt is soluble.
Write the dissociation equation of soluble and insoluble salts.</td><td></td></tr>
<tr><td>4.1.8
4.1.9

4.1.10</td><td>Describe an acid or base as strong or weak memorizing the strong acids and bases.
Describe that acids are typically polar substances so they are generally soluble in water.
Write the dissociation equation of strong and weak acids/bases.</td><td></td></tr>
</table>

Topic	Objectives		I Can
	4.1.11	Predict how an increase of temperature affects solubility.	
	4.1.12	Write solubility (the max amount of salt that can dissolve) in terms of g solute/100 g water.	
	4.1.13	Use a solubility curve to determine the mass of solute in a saturated solution at a given temp.	
	4.1.14	Predict how an increase of pressure affects solubility of a gaseous solute.	
4.2 Concentration	4.2.1	Calculate the moles of solute, volume of solution, or concentration of a solution in molarity using the definition of molarity ($\underline{M} = \frac{\text{moles solute}}{\text{liters solution}}$).	
	4.2.2	Compare the concentration of two solutions quantitatively and conceptually.	
	4.2.3	Describe how to prepare a solution with a given concentration and volume.	
	4.2.4	Use the dilution formula ($\underline{M}_{conc}V_{conc} = \underline{M}_{dil}V_{dil}$) to calculate the concentration or volume of a stock or diluted solution.	
	4.2.5	Describe how to dilute a solution to given concentration.	
4.3 Properties of Solutions	4.3.1	Provide examples of and describe common properties of acids and bases (e.g. sour/bitter, corrosive/caustic)	
	4.3.2	Determine whether a substance is a strong, weak, or non-electrolyte in an aqueous solution.	
	4.3.3	Recall the Arrhenius and Brønsted-Lowry definitions of acids and bases.	
	4.3.4	Identify the acid, base, conjugate acid, and conjugate base in a Brønsted-Lowry acid-base reaction.	
	4.3.5	Write the conjugate base given an acid, and vice versa.	
	4.3.6	Explain the degree to which water autoionizes using its K_w.	
	4.3.7	Find the pH, pOH, $[H^+]$, and $[OH^-]$ of a solution.	
	4.3.8	Determine whether a solution is acidic, neutral, or basic/alkaline using an indicator (e.g. cabbage juice, phenolphthaelin, or litmus) or by its pH/pOH/$[H^+]$/$[OH^-]$.	
4.4 Describing Solutions	4.4.1	Find the concentration of all dissolved species and the pH for a strong acid or a strong base.	
	4.4.2	Write the expression for the K_a (for a weak acid), K_b (for a weak base), or K_{sp} (for an insoluble salt).	
	4.4.3	Describe weak acid/base solutions as an equilibrium between molecular compounds and ionized particles and write the equation for the physical process.	
	4.4.4	Compare the strength of weak acids or bases using their K_a or K_b.	
	4.4.5	Find the concentration of all dissolved species, the pH, and percent ionization for a weak acid or a weak base given the K_a or K_b.	
	4.4.6	Find the concentration of all dissolved species for a soluble salt solution.	
	4.4.7	Find the concentration of all dissolved species and the solubility for an insoluble salt solution given the K_{sp}.	
Topic 5 – Chemical Reactions			
5.1 Types of Reactions	5.1.1	Categorize a reaction as double replacement, single replacement, combination, decomposition, or combustion.	
	5.1.2	Balance an equation by inspection	
	5.1.3	Write the balanced equations given the reactants for the synthesis of a binary compound, decomposition of a binary compound, and combustion of an organic compound.	
5.2 Reactions in Aqueous Solutions	5.2.1	Write the net ionic equation of a reaction in solution.	
	5.2.2	Write the balanced molecular and net ionic equations of a double replacement reaction.	
	5.2.3	Identify the products formed in the double replacement reaction. (precipitate, gas, weak acid or water)	

Topic	Objectives		I Can
5.3 Redox Reactions	5.3.1	Predict the products of a single replacement reaction given reactants.	
	5.3.2	Identify the oxidation states (charges) of elements and monoatomic ions.	
	5.3.3	Describe a redox reaction in terms of losing and gaining electrons.	
	5.3.4	Identify whether a reaction is redox process.	
	5.3.5	Describe in a single replacement reaction which elements are oxidized or reduced, and which substances are the reducing and oxidizing agents.	
	5.3.6	Write the half-reactions of a single replacement reaction.	
5.4 Stoichiometry	5.4.1	Perform one-step stoichiometry calculations to find moles of product formed.	
	5.4.2	Perform multi-step stoichiometry calculations to find the yield in moles, molecules, grams, and liters (of gas at STP).	
5.5 Limiting Reactants	5.5.1	Determine which reactant is limiting and the amounts of the substances in the final mixture, given amounts of the starting material.	
	5.5.2	Determine the percent yield, given the actual yield.	
Topic 6 – Thermochemistry			
6.1 Heat	6.1.1	Determine whether a process is endothermic or exothermic in terms of direction of heat flow, chemical equation, breaking or forming bonds, potential energy graph, sign of q.	
	6.1.2	Calculate the mass, change in temperature, specific heat, or heat transfer of an object using the calorimetry equation. ($q = m \times C \times \Delta T$)	
	6.1.3	Compare the specific heat of two objects quantitatively and conceptually.	
6.2 Changes in States	6.2.1	Determine whether energy is absorbed or released when a substance is changing states.	
	6.2.2	Calculate the energy required or released during a change in state using the heats of fusion and vaporization. ($q = n \times \Delta H$)	
	6.2.3	Compare the heats of formation and vaporization quantitatively and conceptually.	
	6.2.4	Draw and describe the heating curve or cooling curve for a substance.	
	6.2.5	Describe what is happening molecularly as heat is added or removed from a substance.	
	6.2.6	Calculate the energy released/required for a two- or three- step process.	
6.3 Enthalpy of Reaction	6.3.1	Use the Enthalpy of Reaction (ΔH_{rxn}) in heat calculations. ($\Delta H_{rxn} = \frac{q}{n_{rxn}}$)	
	6.3.2	Calculate the Enthalpy of Reaction (ΔH_{rxn}) using Enthalpy of Formation (ΔH_f) values. $\Delta H_{rxn} = \sum_{products} n\,\Delta H_f - \sum_{reactants} n\,\Delta H_f$	
	6.3.3	Calculate the Enthalpy of Reaction (ΔH_{rxn}) using Bond Energies.	
Topic 7 – Kinetics			
7.1 Collision Theory	7.1.1	Use collision theory to explain factors that determine the rate of a chemical reaction.	
	7.1.2	Explain how surface area of solid, pressure of gas, concentration of solution, temperature, and presence of a catalyst affects the rate of a chemical reaction.	
	7.1.3	Identify the energy of reactants, energy of products, change of energy, and activation energy in a potential energy graph.	
	7.1.4	Describe how a change in temperature affects the rate of reaction using the potential energy and kinetic energy graphs.	
	7.1.5	Describe how the presence of a catalyst affects the rate of reaction using the potential energy and kinetic energy graphs.	
	7.1.6	Describe how a catalyst works to speed up the rate of a reaction.	
	7.1.7	Identify the intermediates and catalysts in a reaction given a proposed mechanism.	
	7.1.8	Identify the rate-limiting step of a proposed mechanism as the slow step, having the largest activation energy.	

Topic	Objectives		I Can
7.2 Reaction Rates	7.2.1	Describe numerically and graphically the rate of a reaction as the change in concentration of a reactant over time.	
	7.2.2	Use the rate of reaction to compare the rates of substances in a reaction given the balanced equation. For a generic reaction $a\text{ A} + b\text{ B} \rightarrow c\text{ C} + d\text{ D}$ Rate of reaction $= -\frac{1}{a}\frac{\Delta[\text{A}]}{\Delta t} = -\frac{1}{b}\frac{\Delta[\text{B}]}{\Delta t} = +\frac{1}{c}\frac{\Delta[\text{C}]}{\Delta t} = +\frac{1}{d}\frac{\Delta[\text{D}]}{\Delta t}$	
	7.2.2	Determine the order with respect to each reactant and the rate law of a reaction given initial rate data.	
	7.2.3	Find the value and Topics for the rate constant given the rate of reaction at particular reactant concentrations.	
Topic 8 – Equilibrium			
8.1 Equilibrium	8.1.1	Explain how reversible reactions can reach equilibrium when rates of forward and reverse reactions are equal.	
	8.1.2	Describe what is observed when a system reaches equilibrium.	
	8.1.3	Write the equilibrium constant of a chemical equation.	
	8.1.4	Describe the relative amounts of substances of a system at equilibrium using the magnitude of the value of the equilibrium constant.	
	8.1.5	Determine whether a reversible reaction proceeds toward the right or the left by comparing the reaction quotient Q to K.	
	8.1.6	Find the equilibrium constant of a reaction when the equation is reversed or a multiple of the original equation.	
8.2 ICE Box Problems	8.2.1	Find the amounts of each substance in a reaction mixture at equilibrium given initial amounts and value of K using an ICE Chart.	
8.3 Le Chatelier's Principle	8.3.1	Predict the direction a reaction will shift and how the amounts of substances in the reaction mixture will change if reactants/products are added/removed, the pressure/volume of a gas sample is changed, or if the temperature is changed.	
Topic 9 – Atomic Structure			
9.1 Atomic Orbitals	9.1.1	Find the wavelength, frequency, speed, and energy of a given EM radiation.	
	9.1.2	Explain through emission spectra that electronic energy levels are quantized.	
	9.1.3	Calculate the difference of energy between two levels given the wavelength or frequency of light emitted.	
	9.1.4	Recognize the basic shapes of s, p, and d orbitals.	
9.2 Electron Configuration	9.2.1	Describe atomic orbitals in a multi-electron atom in terms of their relative energies and degeneracy.	
	9.2.2	Relate the position of at atom on the periodic table and its predicted electron configuration.	
	9.2.3	Write the electron configuration of a neutral atom in long form and short form.	
	9.2.4	Write the electron configuration of an ion in long form and short form.	
	9.2.5	Identify the number of valence electrons in a given atom or ion.	
	9.2.6	Explain the similarities in chemical properties, and common monoatomic ions formed using their electron configurations.	

Topic	Objectives		I Can
9.3 Periodic Trends	9.3.1	Describe and explain the trend of effective nuclear charge between elements across a period and of shielding by core electrons down a family.	
	9.3.2	Describe and explain the trends of atomic radius, ionization energy, and electronegativity across a period and down a family in the periodic table.	
	9.3.3	Explain the exceptions to the trends in ionization going across a period (e.g. Be and B, N and O).	
	9.3.4	Determine the number of valence electrons in an unknown element given the successive ionization energies.	
	9.3.5	Compare atomic radius, ionization energy, and electronegativity between two atoms across a period and down a family in the periodic table.	
	9.3.6	Compare the atomic and ionic radii of an element, and the ionic radii of ions with same number of valence electrons.	
Topic 10 – Molecular Structure			
10.1 Types of Bonding	10.1.1	Describe a bond as metallic, non-polar covalent, polar covalent, or ionic based on the difference in electronegativity between the two bonded atoms.	
	10.1.2	Explain the role of electrons in a metallic, non-polar covalent, polar covalent, or ionic of bond formation.	
	10.1.3	Draw the Lewis structures of neutral atoms and ions.	
	10.1.4	Draw the Lewis structure representation of an ionic compound.	
	10.1.5	Qualitatively compare the strength of ionic bonds using Coulomb's Law.	
10.2 Lewis Structures	10.2.1	Draw the Lewis dot structure of covalent compounds or ions using the octet rule.	
	10.2.2	Draw the Lewis dot structure of covalent compounds or ions containing double and triple bonds.	
	10.2.3	Draw the Lewis dot structures of covalent compounds or ions that display resonance.	
	10.2.4	Draw the Lewis dot structures of covalent compounds or ions that are contain central atoms that are hypervalent (e.g. S, P) or hypovalent (e.g. B).	
	10.2.5	Identify and describe the sigma and pi bonds in a covalent compound.	
	10.2.6	Compare the bond energy and bond length of covalent bonds between two elements.	
10.3 Molecular Shapes	10.3.1	Identify the steric number of an atom in a compound, determine the hybridization on that atom, and predict the bond angles of the electron domains around that atom.	
	10.3.2	Predict the overall geometry of the electron domains and between bonded atoms of a covalent compound.	
	10.3.3	Describe a covalent compound as a polar or non-polar molecule.	
10.4 Inter-Molecular Forces	10.4.1	Identify the intermolecular forces between two covalent compounds.	
	10.4.2	Classify a solid as ionic, metallic, network covalent, and molecular, and describe properties of each type of solid.	
	10.4.3	Determine what type of intermolecular forces are involved in a molecular solid and its properties.	
Topic 11 – Properties of Gases			
11.1 Kinetic Molecular Theory	11.1.1	Describe the properties of gases in terms of the kinetic molecular theory.	
	11.1.2	Compare the average speed of gas molecule at a particular temperature.	
	11.1.3	Convert between different units of pressure (atm, mmHg, torr, kPa, psi), temperature (K and C), and volume (L and mL).	
11.2 Gas Laws	11.2.1	Describe how pressure, volume, and temperature are related using Boyle's, Charles's, Gay Lussac's, and Combined Gas Laws qualitatively and graphically.	
	11.2.2	Calculate the properties of a sample of gas using the gas laws (Boyle's, Charles's, Gay-Lussac's, and Combined Gas Laws) when pressure, volume and/or temperature is/are changed.	
	11.2.3	Describe the properties of a gas sample (pressure, volume, temperature, number of moles, mass, density) using the Ideal Gas Law.	
11.3 Gas Mixtures	11.3.1	Describe the properties of a mixture of gases using Dalton's Law of Partial Pressures.	

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AP Chemistry

Name ____________________________

Preparation for AP Chemistry

Period ________ Date _______________

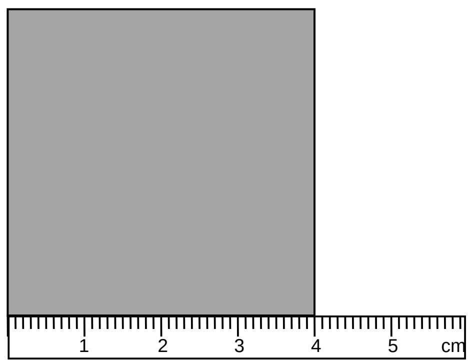
R E V I E W   P R O B L E M S

The following are problems that students entering AP Chemistry are expected to solve and answer without difficulty. You may use a scientific calculator. A periodic table and other helpful information are provided on the last page. If you are finding the need to refer to a textbook or other resources (e.g. notes or a tutor), or are spending an enormous amount of time on any problem, please review that topic thoroughly.

Answer any explanation questions briefly with complete sentences. Complete work for computational problems should be shown, including with I.E.S.A. (if using a formula), all units shown, and answer with the correct number of significant digits.

Topic 1 – Foundations of Chemistry

1. The side of a 169.3 g aluminum cube was measured by Group A with the ruler shown on the right.
   a. Record the measurement of the side of the cube. Circle the uncertain digit. Explain why that is the uncertain digit.



   b. Calculate the density of the aluminum cube. (Volume = $(\text{side})^3$)

   c. Convert the density calculated in part c to units of lb/ft^3.
   (1 kg = 2.2046 lb, 1 ft = 12 in, 1 in = 2.54 cm)

   d. If the accepted density of aluminum is 167 lb/ft^3, calculate the percent error of your calculated density.

   e. Group B determined the density of the aluminum block to be 164 lb/ft^3. Describe the accuracy and precision of this calculation, compared to that of Group A.

2. Suppose the figure in problem 1 shows the measurement of a square sheet of aluminum foil with a mass of 1.90×10^{-3} lb. Find the thickness of the aluminum foil in mm. Write your answer in scientific notation.
(Use density of aluminum = 167 lb/ft^3 from part 1 part d.)

Topic 2 – Elements and Compounds

3. Complete the following table:

Isotope	Atomic Number	Mass Number	Protons	Neutrons	Electrons	Charge
$^{63}Ni^{2+}$						
		32		17		–3
		1			0	
			38	51	36	

4. On the periodic table below, identify the:
- alkali metals
- alkaline earth metals
- halogens
- inner-transition metals
- noble gases
- other metals
- other non-metals
- semi-metals
- transition metals

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	71 Lu	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	103 Lr	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn						
				57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb
				89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No

5. Of the elements on the periodic table, list:
- The diatomic elements. (7)
- The elements that naturally exist as liquids: (2)
- The elements that naturally exist as gases: (11)

6. There are two naturally occurring isotopes of gallium: Ga-69 (mass = 68.92557 amu) and Ga-71 (mass = 70.92470 amu). If the atomic mass of gallium is 69.723 amu, find the percent abundance of each isotope.

7. Write the formulas for the following ions. (These need to be memorized.)

hypobromite ________	hydrogen ________	dichromate ________
iodate ________	thiocyanate ________	nitrate ________
permanganate ________	perbromate ________	bromate ________
sodium ________	nitrite ________	bicarbonate ________
ammonium ________	hypoiodite ________	chromate ________
bisulfate ________	nickel ________	hydronium ________
phosphite ________	chlorate ________	peroxide ________
iodite ________	stannic ________	ferric ________
silver ________	oxide ________	chloride ________
cuprous ________	fluoride ________	strontium ________
phosphate ________	cyanide ________	periodate ________
iodide ________	thiosulfate ________	potassium ________
sulfite ________	hydroxide ________	chlorite ________
lithium ________	mercuric ________	ferrous ________
calcium ________	cupric ________	bromide ________
sulfide ________	hypochlorite ________	perchlorate ________
acetate ________	plumbic ________	sulfate ________
mercurous ________	stannous ________	zinc ________
bromite ________	plumbous ________	barium ________
aluminum ________	carbonate ________	magnesium ________

8. For each of the following compounds, classify it as an Ionic Compound, Covalent Compound, or Acid. Then write its name.

Formula	**Type**	**Name**
H_2SO_4		
$LiNO_3$		
Na_2S		
Fe_2O_3		
$Pb(BrO_3)_2$		
P_4O_{10}		
$BaBr_2$		
HBr		
CaO		
$Zn(C_2H_3O_2)_2$		
HNO_2		
CuI		
N_2O		
$KClO_3$		
H_3PO_3		
PCl_5		
$HBrO_4$		
$Al_2(CO_3)_3$		
$Na_2S_2O_3$		
Hg_2Cl_2		
HI		
SO_3		
Cs_2O_2		
$Sn_3(PO_3)_2$		

9. For each of the following compounds, classify it as an Ionic Compound, Covalent Compound, or Acid. Then write its formula.

Name	**Type**	**Formula**
hydrofluoric acid		
silver nitrate		
cupric cyanide		
aluminum oxide		
potassium dichromate		
sulfurous acid		
dinitrogen tetroxide		
carbonic acid		
hypobromous acid		
manganese(IV) oxide		
cobalt(III) sulfate		
lithium bicarbonate		
stannic oxide		
perioidic acid		
iodine heptafluoride		
oxygen difluoride		
rubidium chloride		
hydrocyanic acid		
plumbic bromide		
ammonium carbonate		
magnesium fluoride		
chromium(III) chloride		
chlorous acid		
potassium iodide		

10. For each of the following compounds, calculate the % *oxygen* by mass of each element in the compound, and its molar mass.

	% Oxygen	Molar Mass
a. diboron trioxide		
b. ammonium nitrate		
c. mercurous phosphate		
d. calcium acetate		

11. Find the following.
    a. How many moles are in a 3.55×10^{24} molecule sample of O_3?
    b. What is the mass of a 12.5 L sample of nitrogen gas at STP?
    c. What is the mass in grams of a 0.72 mol sample of sulfur dioxide?
    d. How many moles are in a 30.5 L sample of carbon dioxide gas at STP?
    e. What is the volume at STP of a sample that contains 4.00×10^{23} molecules of methane, CH_4?
    f. How many molecules are in an 18.2 L sample of helium gas at STP?
    g. How many molecules are in a 1.20 mol sample of oxygen gas? How many oxygen atoms are in this sample?
    h. How many moles are in a 75.4 g sample of sulfur tetrachloride?
    i. What is the volume at STP of a 102 g sample of propane gas, C_3H_8?
    j. What is the mass of a sample that contains 1.22×10^{22} argon atoms?
    k. What is the volume in liters of a 0.14 mol sample of chlorine gas at STP?
    l. How many molecules are in a 65.0 g sample of water, H_2O?

12. A 30.15 g sample of a gas containing only N and O occupies 14.7 L at STP.
    a. Find the molar mass of this compound.
    b. If the compound is 30.4% N by mass, find the empirical and molecular formulas for this compound.
    c. How many oxygen atoms are in this sample?
    d. Write the balanced chemical equation for the formation of this compound. Include states.

13. An 80.50 g sample of an unknown oxide of chromium (Cr_xO_y) is heated until all of the oxygen is vaporized, leaving 55.08 g of solid chromium metal. The molar mass of this unknown compound is between 150 and 160 g/mol.
    a. Find the % composition of each element of the unknown compound.
    b. Find the empirical and molecular formulas of the unknown compound.
    c. Write the balanced chemical equation for this decomposition. Include states.
    d. Find the volume at STP of the oxygen gas that is vaporized.
    e. Find the number of chromium atoms in this sample.

Topic 3 – Organic Chemistry

14. Draw the structural formulas for and name the isomers for following compounds:
    a. Nine isomers for $C_4H_6Cl_2$.

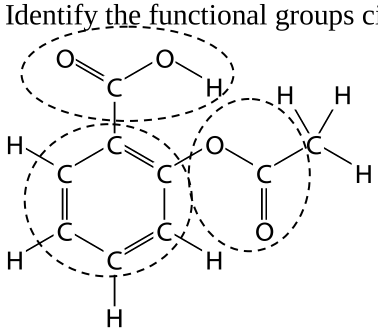
b. Three isomers for $C_2H_2Br_2$.

c. Two isomers for C_4H_6.

15. Write the general formulas for each of the following functional groups.

Aldehyde	Alcohol	Ketone	Ether
Ester	Carboxylic acid	Amine	

16. Identify the functional groups circled in the compound below:



17. Write the balanced equation for the combustion of butanone, C_4H_8O.

Topic 4 – Salts and Solutions

18. Complete the table below. Each box may contain one, more than one, or no substances.

a. Predict the color in cabbage juice and conductivity with a light bulb of each of the following substances in the table below.

AgCl	$C_{12}H_{22}O_{11}$	HCl	$HC_2H_3O_2$
NaCl	NaOH	NH_3	

b. Categorize each box with one or more of the following:

strong acid	weak acid	strong base	weak base
soluble salt	insoluble salt	molecular solid	

		Electrical Conductivity with a Light Bulb		
		Bright	**Dim**	**Dark**
Cabbage Juice Color	**Red**			
	Purple			
	Green			

c. Write the dissociation equation in each box for each of the substances. If the substance does not dissociate, say so.

	Dissociation Equation
AgCl	
$C_{12}H_{22}O_{11}$	
HCl	
$HC_2H_3O_2$	
NaCl	
NaOH	
NH_3	

19. Using the Solubility Rules, determine which of the following are soluble in aqueous solution, and write their dissociation equations.

		Soluble?	Dissociation Equation
a.	$Al(C_2H_3O_2)_3$	Soluble | Insoluble	______________________
b.	$Al(OH)_3$	Soluble | Insoluble	______________________
c.	$Ba(OH)_2$	Soluble | Insoluble	______________________
d.	$Ca(ClO_3)_2$	Soluble | Insoluble	______________________
e.	$CaSO_4$	Soluble | Insoluble	______________________
f.	Hg_2Cl_2	Soluble | Insoluble	______________________
g.	K_2CrO_4	Soluble | Insoluble	______________________
h.	$Mg_3(PO_4)_2$	Soluble | Insoluble	______________________
i.	$SnCl_2$	Soluble | Insoluble	______________________
j.	$ZnSO_4$	Soluble | Insoluble	______________________

20. Find each of the following.

a. What is the concentration of a 1.50 L solution with 0.300 mol HCl?

b. What is the concentration when 42.6 g $CaBr_2$ is dissolved in 800 mL of water?

c. How many moles of KF are in 200 mL of a 0.400 M solution?

d. What is the mass required to prepare 500 mL of a 1.30 M solution of $NaNO_3$?

e. What is the volume of a 3.00 M solution containing 0.850 mol $Mg(ClO_3)_2$?

f. What is the volume of a 0.250 M solution in which 30.0 g $CaCl_2$ is dissolved?

g. What volume of a 3.00 M solution of HCl is required to prepare 200.0 mL of a 0.120 M solution?

h. What is the concentration of the solution when 300 mL of a 1.80 M NaI solution is diluted to 4.00 L?

i. What is the final volume of a 0.600 M KNO_3 solution prepared with 50.0 mL of a 1.50 M solution?

j. What is the concentration of a stock solution of $HC_2H_3O_2$ from which 20.0 mL was used to prepare a 200 mL sample of a 0.050 M solution?

21. Complete the following table:

	$[H^+]$	$[OH^-]$	pH	pOH	Acid/Base/Neutral
a.	1.0×10^{-5}				
b.		0.0500			
c.			2.00		
d.				–0.10	

22. Find the pH of the following solutions:

a. 0.050 M solution of $Ca(OH)_2$

b. 0.050 M solution of HCl.

c. 0.050 M solution of $HC_2H_3O_2$. The K_a of $HC_2H_3O_2$ is 1.8×10^{-5}.

Topic 5 – Chemical Reactions

23. For each of the following reactions, identify the type of reaction (i.e. combustion, decomposition, double replacement, single replacement, or synthesis), write the balanced molecular equation, and write the net ionic equation. Include states when possible.

For reactions e, f, g, h, I, and j, identify the element that is oxidized and the element that is reduced and how the oxidation changes for each.

a. Solutions of potassium chromate and silver nitrate are mixed.

Type:

Balanced Molecular Equation

Net Ionic Equation

b. Propane gas is burned in oxygen.

Type:

Balanced Molecular Equation

Net Ionic Equation

c. A gaseous sample of butane is burned in air.

Type:

Balanced Molecular Equation

Net Ionic Equation

d. Solutions of lead(II) acetate and hydrochloric acid are mixed.

Type:

Balanced Molecular Equation

Net Ionic Equation

e. Aluminum metal reacts completely with chlorine gas.

Type:

Balanced Molecular Equation

Net Ionic Equation

Element Oxidized: Element Reduced:

f. A piece of aluminum foil is dipped in a solution of cupric chloride.

Type:

Balanced Molecular Equation

Net Ionic Equation

Element Oxidized: Element Reduced:

g. Silver metal is combined with oxygen gas.

Type:

Balanced Molecular Equation

Net Ionic Equation

Element Oxidized: Element Reduced:

h. A piece of zinc metal is placed in a solution of hydrochloric acid.

Type:

Balanced Molecular Equation

Net Ionic Equation

Element Oxidized: Element Reduced:

i. Water is electrolyzed.

Type:

Balanced Molecular Equation

Net Ionic Equation

Element Oxidized: Element Reduced:

j. Solid cupric oxide is strongly heated.

Type:

Balanced Molecular Equation

Net Ionic Equation

Element Oxidized: Element Reduced:

24. Determine which pairs of the following solutions, when mixed, form precipitates. Write the balanced molecular and balanced net ionic equations for these mixtures. $AgNO_3$ K_2SO_4 Na_2CrO_4 $CaCl_2$

Balanced Molecular Equation:

Net Ionic Equation:

Balanced Molecular Equation:

Net Ionic Equation:

Balanced Molecular Equation:

Net Ionic Equation:

Balanced Molecular Equation:

Net Ionic Equation:

25. Consider the combustion of 30.0 g of propane in excess oxygen.
    a. Write the balanced molecular equation for this reaction.
    b. Find the maximum mass of H_2O that can form.
    c. Find the volume of CO_2 produced at STP.

26. A 58.0 g sample of $KClO_3$ (s) decomposed to form KCl (s) and O_2 (g).
    a. Write the balanced molecular equation for this reaction.
    b. Find the theoretical yield in L of O_2 (g).
    c. If 14.5 L of O_2 (g) was produced at STP, find the percent yield.

27. Consider 1.35 g Mg placed in 100.0 mL of a 1.50 M HCl solution.
    a. Write a balanced molecular equation for this reaction.
    b. Find the maximum volume of H_2 (g) that can be produced at STP.
    c. If 0.950 L of H_2 was produced at STP, find the percent yield.

28. In the reaction $8\ P_4\ (s) + 3\ S_8\ (s) \rightarrow 8\ P_4S_3\ (s)$, 20.0 g of P_4 and 40.0 g of S_8 are combined.
    a. What is the theoretical yield of P_4S_3? What are the limiting and excess reactants?
    b. What is the mass of the excess reactant that remains if the reaction goes to completion?
    c. If 30.0 g of P_4S_3 are actually produced, find the percent yield.

Topic 6 – Thermochemistry

29. Blocks A and B both have the same mass and are at 20°C, but are made of different metals.
    a. If blocks A and B are heated with the 4500 J of energy, the temperatures of the blocks are 25°C and 40°C, respectively. Which block has the greater specific heat? Explain briefly.
    b. If the specific heat of block A is larger than that of block B, which would require more energy to raise the temperature to 50°C? Explain briefly.
    c. If the specific heat of block A is larger than that of block B, which would end up at a higher temperature when 3000 J is released? Explain briefly.

30. Find the following.
    a. What is the mass of a sample of hexane if 500 J of energy increases the temperature from 15.1°C to 17.6°C? The specific heat of hexane is 2.26 J/g·°C.
    b. What is the final temperature when 4500 J of energy is released from a 245 g piece of glass at 35°C? The specific heat of glass is 0.84 J/g·°C.
    c. How much energy is required to heat up a 45.8 g of copper from 11.5°C to 26.1°C? The specific heat of copper is 0.39 J/g·°C.
    d. What is the final temperature when a 32.0 g aluminum block that is 100°C is placed in 180.0 mL water at 22.4°C? The specific heats of aluminum and water are 0.900 J/g·°C and 4.18 J/g·°C. The density of water is 1.00 g/mL.

31. Sketch the part of the heating curve corresponding to each problem. Use the data below for water to find each of the following. ΔH_{fus} = 6.01 kJ/mol, ΔH_{vap} = 40.68 kJ/mol, C_{ice} = 2.10 J/ g·°C, C_{water} = 4.18 J/ g·°C, C_{steam} = 2.08 J/g·°C.

a. 50.0 g water at 80.0°C is vaporized completely to 100.0°C.

b. 4.50 g ice at –15.0°C is heated to 40.0°C.

c. 30.0 g vapor at 105°C is cooled to 50.0°C.

d. 15.0 g ice at 0°C is heated to 20°C.

32. What is the ΔH_{rxn} for the following reaction if 70.4 kJ of energy is absorbed when 28.2 L of NO (g) is produced at STP?

$$2\ NO_2\ (g) \rightarrow 2\ NO\ (g) + O_2\ (g)$$

33. Given the following reaction:

Fe_2O_3 (s) + 3 CO (g) → 2 Fe (s) + 3 CO_2 (g) ΔH_{rxn} = –26.74 kJ/mol

a. How many grams of Fe are produced if 50.8 kJ of energy is released?

b. How much energy is released if 50.2 L CO (g) at STP are used?

c. What is the ΔH_f for Fe_2O_3 (s) if ΔH_f of CO and CO_2 are –110.5 kJ/mol and –393.5 kJ/mol, respectively?

34. Given the following reaction: 2 O_3 (g) → 3 O_2 (g) ΔH_{rxn} = –286 kJ/mol

a. What is the ΔH_{rxn} of the reaction 3 O_2 (g) → 2 O_3 (g)?

b. What is the ΔH_f of O_3?

c. The bond energy of O_2 (O=O) is 495 kJ/mol. What is the bond energy of the O--O bond in O_3 (O--O--O)?

Topic 7 – Kinetics

35. Answer the following using ideas of Collision Theory.

a. State the main ideas of Collision Theory:

For a reaction to occur, molecules must ___________ with enough

______________ in the correct _________.

b. State five ways to increase the rate of a reaction.

-
-
-
-
-

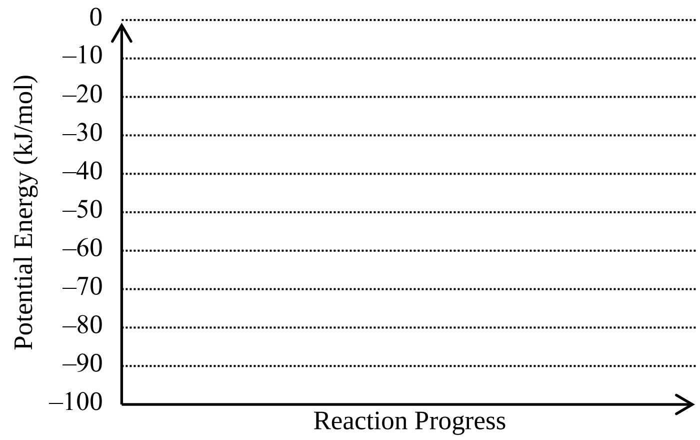
c. The temperature of a system is related to the __________ _________

__________ of the molecules.

d. A catalyst ________ the rate of reaction by: _____________________

_____________________.

36. Sketch the Potential Energy graph and find the values.



$A \leftrightarrows B$

- $\Delta H_{f,A}$ =
- $\Delta H_{f,B}$ = –50 kJ/mol

$\Delta H_{f,T.S.\ uncat}$ =

$\Delta H_{f,T.S.\ cat}$ =

Forward Reaction:

- $E_{a,uncat}$ =
- $E_{a,cat}$ = 40 kJ/mol
- ΔH_{rxn} = +20 kJ/mol
- [endo | exo]

Reverse Reaction:

- $E_{a,uncat}$ = 30 kJ/mol
- $E_{a,cat}$ =
- ΔH_{rxn} =
- [endo | exo]

37. Given the following reaction mechanism:

Step 1: $H_2O_2 + 2\ Br^- + 2\ H^+ \rightarrow 2\ H_2O + Br_2$ fast

Step 2: $H_2O_2 + Br_2 \rightarrow O_2 + 2\ H^+ + 2\ Br^-$ slow

a. Write the equation for the overall reaction.

b. Identify any intermediates and catalysts.

Intermediates:

Catalysts:

38. The following data was collected for the reaction
$2\ NO\ (g) + O_2\ (g) \rightarrow 2\ NO_2\ (g)$.

Time (s)	[NO] (M)
0.00	0.10
0.50	0.078
1.00	0.061
1.50	0.047

a. What is the average rate of disappearance of NO between t = 0.50 s and t = 1.00 s?

b. What is the average rate of appearance of NO_2 between t = 1.00 s and t = 1.50 s?

39. Initial rate data was collected for the reaction
$2\ NO\ (g) + O_2\ (g) \rightarrow 2\ NO_2\ (g)$

Trial	[NO] (M)	[O_2] (M)	Rate of disappearance of NO (M/s)
1	0.100	0.100	4.40×10^{-2}
2	0.200	0.100	8.80×10^{-2}
3	0.100	0.300	1.32×10^{-1}

a. Find the rate law for the reaction.

b. Find the value of the rate constant, k. Include units.

40. Initial rate data was collected for the reaction
$O_2\ (g) + 2\ Cl_2\ (g) \rightarrow 2OCl_2\ (g)$

Trial	[O_2] (M)	[Cl_2] (M)	Rate of appearance of OCl_2 (M/s)
1	0.100	0.100	0.120
2	0.400	0.100	0.120
3	0.400	0.200	0.480

a. Find the rate law for the reaction.

b. Find the value of the rate constant, k. Include units.

Topic 8 – Equilibrium

41. Write the equilibrium constant expression and find the value for the following reactions. Determine whether the reaction is reactant-favored or product-favored.

a. $2\ NH_3\ (g) \leftrightarrows N_2\ (g) + 3\ H_2\ (g)$
$[NH_3]_{eq} = 5.0 \times 10^{-5}$ M, $[N_2]_{eq} = 0.096$ M, $[H_2]_{eq} = 0.29$ M

b. $PbCl_2\ (s) \leftrightarrows Pb^{2+}\ (aq) + 2\ Cl^-\ (aq)$
$[Pb^{2+}]_{eq} = 0.016$ M, $[Cl^-]_{eq} = 0.032$ M

42. Given the reaction $2\ NO_2\ (g) \leftrightarrows N_2O_4\ (g)$

a. Write the equilibrium constant expression for the reaction.

b. If only 0.500 mol of NO_2 (g) is placed in a 1.0 L container and the reaction is allowed to come to equilibrium, 0.186 mol of N_2O_4 (g) is formed. Find the value of K_{eq}.

c. Using the value of K_{eq} from part a, if $[NO_2]$ = 0.090 M and $[N_2O_4]$ = 0.205 M, is the reaction at equilibrium? If not, in which direction will it shift to reach equilibrium?

d. Using the value of K_{eq} from part a, if $[NO_2]$ = 0.138 M and $[N_2O_4]$ = 0.181 M, is the reaction at equilibrium? If not, in which direction will it shift to reach equilibrium?

e. Using the value of K_{eq} from part a, in a reaction at equilibrium, $[N_2O_4]_{eq}$ = 0.102 M. Find $[NO_2]_{eq}$.

43. Given the reaction $H_2\ (g) + I_2\ (g) \leftrightarrows 2\ HI\ (g)$ $\quad K_{eq} = 45.1$

a. Find the concentrations of all substances at equilibrium if 1.50 M of H_2 (g) and 1.50 M of I_2 (g) are initially placed in an empty container.

b. Find the concentrations of all substances at equilibrium if 0.20 M of H_2 (g), 0.20 M of I_2 (g), and 4.0 M of HI (g) are initially placed in an empty container.

44. Consider the reaction *cis*-2-butene $\leftrightarrows$ *trans*-2-butene $\quad K_{eq} = 4.0 \times 10^{-3}$

a. If 0.15 M *cis*-2-butene is placed in a container, find the concentrations of all substances at equilibrium.

b. If 0.20 M of both isomers are placed in a container, find the concentrations of all substances at equilibrium.

45. Consider a system at equilibrium that undergoes the following reaction:

$$2\ SO_2\ (g) + O_2\ (g) \leftrightarrows 2\ SO_3\ (g) + \text{heat}$$

Determine the direction in which the reaction shifts when the following changes were made.

a. Some SO_3 was added to the system. left | no shift | right
b. Some O_2 was removed. left | no shift | right
c. Some SO_2 is added to the system. left | no shift | right
d. The temperature is increased. left | no shift | right
e. A catalyst was added. left | no shift | right
f. The pressure was decreased by increasing the volume. left | no shift | right

Topic 9 – Atomic Structure

46. How many …

a. Electrons can fill all the orbitals in the 5th shell (n = 5). _______
b. Orbitals are in the 4f subshell. _______
c. Half-filled orbitals are in the ground state electron configuration for Se. _______

47. Write the short form and long form ground state electron configurations for the following species. Identify the number of valance electrons for each specie.

	Long Form	Short Form	Val e^-
a. Ni			
b. Al			
c. Ar			
d. As			
e. Au			
f. Zn^{2+}			
g. Ca^{2+}			
h. Br^-			

48. For each of the following, determine the number of valence electrons, and the ion it is likely to form.

	Val e^-	Ion		Val e^-	Ion
a. Sb			b. S		
c. Cs			d. Ga		
e. At					

49. Draw the Lewis Structure for the following species.

a. Br	b. K^+
c. Ca	d. P
e. F^-	f. S^{2-}
g. He	h. Si

50. Rank the following and explain.

a. Increasing atomic radius: P, S, As

b. Increasing ionization energy: K, Ca, Rb

c. Increasing radius: Cl, Cl^{-}, Cl^{2-}

d. Increasing radius: Cl^{-}, Ar, K^{+}

e. Increasing ionization energy: Ca, Ca^{+}, Ca^{2+}

51. The successive ionization energies for unknown element X is listed below. Write the equation that represents each ionization energy, and determine the stable ion that element X forms.

IE_1 = 577 kJ/mol	Equation
IE_2 = 1820 kJ/mol	Equation
IE_3 = 2740 kJ/mol	Equation
IE_4 = 11,600 kJ/mol	Equation
Stable Ion	

Topic 10 – Molecular Structure

52. Identify the type of bonding (nonpolar covalent, polar covalent, ionic, metallic) between the following elements.

a. Ca–P _______________
b. O–H _______________
c. Cu–Sn _______________
d. C–H _______________
e. N–O _______________
f. S–F _______________

53. For each of the following compounds, draw the Lewis Dot Structure. Then describe the shape, identify the bond angles, and determine whether the molecule is polar for compounds a, b, c, d, g, h, j, and l.

	Lewis Structure	Shape
a. CH_2Cl_2		Shape: Bond Angles: Polar?
b. BF_3		Shape: Bond Angles: Polar?

	Lewis Structure	Shape
c. OF_2		Shape: Bond Angles: Polar?
d. NBr_3		Shape: Bond Angles: Polar?
e. PCl_5		Shape: Bond Angles: Polar?
f. SF_4		Shape: Bond Angles: Polar?
g. CO_2		Shape: Bond Angles: Polar?
h. NH_4^+		Shape: Bond Angles: Polar?
i. CN^-		Shape: Bond Angles: Polar?

	Lewis Structure	Shape
j. NO_3^-		Shape: Bond Angles: Polar?
k. C_2H_4		Shape: Bond Angles: Polar?
l. CO		Shape: Bond Angles: Polar?

54. Rank the following. Explain briefly.

a. In increasing bond energy: N_2, O_2, F_2

b. In increasing C–C bond length: C_2H_6, C_2H_4, C_2H_2

c. In increasing C–O bond energy: CH_3OH, CO, CO_2

55. Rank the following in order of increasing melting point. Explain.

a. H_2, N_2, Cl_2

b. KBr, MgS, NaCl

c. HF, HCl, HBr

d. CO_2, CO, CH_3OH

56. Classify each of the following solids (i.e. molecular solid, metallic solid, ionic solid, network covalent solid), and identify the inter-particle forces involved.

Substance	Type of Solid	Inter-Particle Forces
a. SiO_2		
b. I_2		
c. Ca		
d. KCl		
e. CO_2		
f. $C_6H_{12}O_6$		
g. Zn		
h. $C_{diamond}$		
i. H_2O		
j. CaO		

Topic 11 – Properties of Gases

57. Convert 3.2 atm into mmHg, torr, and kPa.

58. The temperature of a sample of N_2 (g) is 10°C. To what temperature (in °C) must the sample be raised to have twice the average kinetic energy?

59. 20.0 g samples of H_2 (g), O_2 (g), and Cl_2 (g) are placed in separate 2.0 L containers at 25°C. Rank the gas samples:

- By increasing molecular speed.
- By increasing average kinetic energy.
- By increasing pressure.

60. Find the following.

a. A sample of SO_2 (g) at 1.4 atm is placed in a 2.0 L container. What is the pressure if the volume is increased to 3.5 L at constant temperature?

b. A sample of Ar (g) at 500 torr is placed in a rigid 10.2 L container at 30°C. To what temperature should the sample be heated in order for the pressure to be raised to 800 torr?

c. A sample of O_2 (g) at 40°C is placed in a balloon with a volume of 1.20 L. If the pressure is held constant, what is the volume of the balloon if the temperature is lowered to 0°C?

d. A sample of CO_2 (g) at 30°C and 98.5 kPa is placed in a 6.0 L container. Find the pressure if the temperature is raised to 60°C and the volume is changed to 3.0 L.

e. At what temperature will a 0.030 mol sample of NO_2 (g) placed in 2.0 L container have a pressure of 1.5 atm?

f. How many grams of He (g) must be placed in a 20.4 L container at 12°C in order for the pressure to be 2.4 atm?

61. A mixture containing 0.010 mol O_3 (g) and 0.080 mol O_2 (g) are placed in a container. What is the total pressure of the container if the partial pressure of O_3 (g) is 0.50 atm?

62. A 40.0 L container contains a mixture of 15.0 g N_2 (g) and 14.0 g Ne (g) at 30°C. Find the partial pressure of each gas in the container.

Periodic Table of the Elements

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 **H** 1.008																	2 **He** 4.003
3 **Li** 6.941	4 **Be** 9.012											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.07	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.87	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.38	31 **Ga** 69.72	32 **Ge** 72.64	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.96	43 **Tc** (98)	44 **Ru** 101.1	45 **Rh** 102.9	46 **Pd** 106.4	47 **Ag** 107.9	48 **Cd** 112.4	49 **In** 114.8	50 **Sn** 118.7	51 **Sb** 121.8	52 **Te** 127.6	53 **I** 126.9	54 **Xe** 131.3
55 **Cs** 132.9	56 **Ba** 137.3	57-71 *	72 **Hf** 178.5	73 **Ta** 180.9	74 **W** 183.8	75 **Re** 186.2	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.1	79 **Au** 197.0	80 **Hg** 200.6	81 **Tl** 204.4	82 **Pb** 207.2	83 **Bi** 209.0	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** (226)	89-103 **	104 **Rf** (265)	105 **Db** (268)	106 **Sg** (271)	107 **Bh** (272)	108 **Hs** (277)	109 **Mt** (276)	110 **Ds** (281)	111 **Rg** (280)	112 **Cn** (285)	113 **Uut** (284)	114 **Fl** (289)	115 **Uup** (288)	116 **Lv** (293)	117 **Uus** (294)	118 **Uuo** (294)

1	
H	← atomic number
1.008	← element symbol
	← atomic mass

57-71 *	57 **La** 138.9	58 **Ce** 140.1	59 **Pr** 140.9	60 **Nd** 144.2	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 152.0	64 **Gd** 157.3	65 **Tb** 158.9	66 **Dy** 162.5	67 **Ho** 164.9	68 **Er** 167.3	69 **Tm** 168.9	70 **Yb** 173.1	71 **Lu** 175.0
89-103 **	89 **Ac** (227)	90 **Th** 232.0	91 **Pa** 231.0	92 **U** 238.0	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)

Solubility Rules for Salts

<u>Always soluble:</u>

- alkali ions, NH_4^+, NO_3^-, ClO_3^-, ClO_4^-, $C_2H_3O_2^-$, HCO_3^-

<u>Generally soluble:</u>

- Cl^-, Br^-, I^- Soluble except w/ Ag^+, Pb^{2+}, Hg_2^{2+}
- F^- Soluble except w/ Pb^{2+}, Ca^{2+}, Ba^{2+}, Sr^{2+}, Mg^{2+}
- SO_4^{2-} Soluble except w/ Pb^{2+}, Ca^{2+}, Ba^{2+}, Sr^{2+}

<u>Generally insoluble:</u>

- O^{2-}, OH^-
  Insoluble except w/ Ca^{2+}, Ba^{2+}, Sr^{2+}, alkali ions, NH_4^+
- CO_3^{2-}, PO_4^{3-}, S^{2-}, SO_3^{2-}, CrO_4^{2-}, $C_2O_4^{2-}$
  Insoluble except with alkali ions, NH_4^+

Strong Acids: HCl, HBr, HI, HNO_3, H_2SO_4, $HClO_3$, $HClO_4$, HIO_4

Strong Bases: LiOH, NaOH, KOH, RbOH, CsOH, $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$

Gases that Form:
$\rightarrow H_2S$ (g)
$\rightarrow H_2CO_3$ (aq) $\rightarrow CO_2$ (g) + H_2O (ℓ)
$\rightarrow H_2SO_3$ (aq) $\rightarrow SO_2$ (g) + H_2O (ℓ)
$\rightarrow NH_4OH$ (aq) $\rightarrow NH_3$ (g) + H_2O (ℓ)