

Chapter One. Matter and Measurements

Classification of Matter. Figure 1.3. Elements and the Periodic Table. Metals, nonmetals and metalloids. Units and unit conversions. Factor-label method. Significant figures and scientific notation. Rounding.

Note 1. A common error when performing long calculations is to round off too soon, thus leading to round-off error. The rule is carry extra significant figures until you get the final answer. Then and only then do you round off to the correct number of significant figures.

Laboratory lore: Be able to calculate “percent difference” between two or more measurements. Be able to calculate “percent error”, comparing your experimental result against a standard or the “true value”.

Define **calorie**. Define **specific heat**. Be able to calculate the specific heat of an unknown (Experiment 7 – Specific heat of a metal). Define **density** and be prepared to explain how you would determine the density of an unknown solid or liquid.

Chapter Two. Atoms and the Periodic Table

Atomic theory. Subatomic particles: proton, neutron, electron, photon. Charge and mass. Atomic number. Isotopes. Periodic Table. Families (groups). Main group elements. S-block, d-block, p-block and f-block elements. Quantum theory. Aufbau – building up the electrons of an element according to the energy level of each. Electron configuration. Subshells, orbitals, spin. Electron-dot symbols. The electromagnetic spectrum.

Families: Define each of the following: alkali metal, alkaline-earth metal, transition metal, halogen, noble gas.

Chapter Three, Ionic Compounds

$M + NM \rightarrow$ Ionic compound (also known as a ‘salt’). Electron transfer. Ionization energy. Properties of ionic compounds that set them apart from molecular compounds. Octet rule. Two systems of naming binary ionic compounds: (1) main-group metal salts; (2) transition metal salts. Complex or polyatomic ions such as ammonium, sulfate, etc.

What does the term ‘formula unit’ refer to.

Chapter Four. Molecular Compounds

$NM + NM \rightarrow$ molecular compound. The covalent bond. Molecular orbital theory. Define bond length. Molecular geometry. Lone pair or non-bonding pair of electrons. Be able to apply VSEPR theory to predict molecular geometry. Electronegativity and the periodic table. The **semi-polar or polar covalent bond**. Molecular symmetry. Permanent dipole moment. Molecules with semipolar or polar covalent bonds may have either a permanent or a zero dipole moment – how can that be? Double and triple bonds. Pi-bonds vs. sigma bonds. Lewis structures (aka Lewis octet structures). The third system of naming: binary molecular compounds. P_2O_5 is diphosphorus pentoxide.

Chapter Five. Classification and Balancing of Chemical Reactions.

The chemical equation must be balanced with respect to mass and charge. Coefficients vs subscripts vs superscripts. The double-displacement reaction: under what 3 or 4 conditions does it go to completion? **Neutralization**. Redox reactions and the concept of **oxidation number**. Know the rules for assigning oxidation numbers. Recognizing redox reactions (study Section 5.7). The **net ionic equation**. Precipitation reactions and the solubility rules.

Redox reaction basics: The oxidized form gets reduced (its oxidation number gets smaller) and is the **oxidizing agent**. The reduced form gets oxidized (its oxidation number gets larger) and is the **reducing agent**.

Chapter Six. Chemical Reactions: Mole and Mass Relationships.

Define **mole**. Avogadro's number, Be able to go from mass to mole and from mole to mass. Units, units, units. Solving mass relationship equations:

Mass of A to moles of A; moles of A to moles of B; moles of B to mass of B.

Remember: mass is conserved but moles are not. **Limiting reagent**. Percent yield.

Formula unit and formula weight – used for ionic compounds, that is, salts, that do not qualify as molecules (see definitions in back of text).

Chapter Seven. Chemical Reaction: Energy, Rates, and Equilibrium

Part One: Energy and chemical bonds. The formation of a chemical bond from the elements is energetically favorable: Examples: $H + H \rightarrow H_2 + \text{energy}$ (covalent bond). $K + Cl \rightarrow KCl + \text{energy}$ (ionic bond). **Bond energy**. Given the bond energies, be able to calculate the energy of a given reaction and tell whether the reaction is **exothermic** or **endothermic**.

Part Two: Chemical thermodynamics. $\Delta G = \Delta H - T\Delta S$. First Law: Energy is conserved. Second Law: All systems tend toward a state of equilibrium or maximum entropy (where entropy is a measure of the disorder of the system). At equilibrium, $\Delta G = 0$ and $\Delta H = T\Delta S$. Spontaneity: study 'important points' at bottom of pg 188 in text.

Kinetics and equilibrium. Reaction rate is the speed at which the reactants disappear or the speed at which the products form and is measured in units of change of concentration per time ($\pm \Delta[\text{reactant or product}]/t$). At equilibrium, rate of forward reaction = rate of reverse reaction.

Given a system at equilibrium, be able to write the expression for the equilibrium constant, K, in terms of the concentrations of reactants and products. **LeChatelier's principle**. (Note: The principle applies to physical – melting, boiling, solubility -- as well as chemical equilibria).

What are the effects of temperature, concentration and catalysts on reaction rates? See Figure 7.4.