AP Chemistry

Chapter 4 Outline

a) General Properties of Aqueous Solutions
   i) Solution = a homogeneous mixture of two or more substances
      (1) Solvent = substance present in greatest quantity
      (2) Solute = the other substance(s) present in a solution
      (3) Aqueous solution = a solution in which the dissolving medium is water
   ii) Electrolytic Properties—tested using "conductivity probe" (Animation)
      (1) Electrolyte = a substance whose aqueous solutions contain ions; ionic substances
         (a) Ionic substances "dissociate" into its component ions as it dissolves
            (i) Water is an effective solvent for ionic compounds because the O is δ- and the H atoms are δ+
            (ii) "solvation" process prevents dissociated cations and anions from recombining, and ions become uniformly dispersed
      (2) Strong electrolytes = solutes that in solution, exist completely or nearly completely as ions; reactions would be written using a single arrow
      (3) Weak electrolytes = solutes that mostly remain as molecules, with only a small fraction ionizing
      (4) Nonelectrolyte = a substance that does not form ions in solution; molecular substances
         (a) Solution consists of intact molecules uniformly dispersed in solvent
         (b) Some molecular species DO ionize in solution: especially ACIDS
            (i) Weak acids form an "equilibrium" between ions and undissociated molecules; use a double headed arrow $\rightleftharpoons$ when writing reactions

b) Precipitation Reactions
   i) Precipitate = an insoluble solid formed by a reaction in solution
   ii) Occur when certain pairs of oppositely charged ions attract each other very strongly
   iii) Solubility = the amount of a substance that can be dissolved in a given quantity of solvent at a certain temperature
      (1) Memorize the key solubility rules (separate sheet)—based on experimental observations
   iv) Exchange (Metathesis) Reactions—(also called "double displacement") Illustration Note the picture model representations!
      (1) General form: $AX + BY \rightarrow AY + BX$
   v) Ionic Equations
      (1) Molecular equation: shows complete chemical formulas of reactants and products
      (2) Complete ionic equation: shows all soluble strong electrolytes as ions
         (a) Spectator ion: ion that appears in identical forms on both sides of equation; present, but plays no direct role
      (3) Net ionic equation: omits all spectator ions; only shows ions & molecules directly involved in reaction VERY IMPORTANT ON AP EXAM

c) Acid-Base Reactions
   i) Acid = substances that ionize in aqueous solutions to form hydrogen ions
      (1) "proton donor"
      (2) hydronium ion, H_3O^+
      (3) monoprotic vs. diprotic acids
         (a) in diprotic acids, only the loss of the first proton results in extensive ionization
   ii) Base = substances that accept (react with) hydrogen ions; produce OH- ions when dissolved in water
      (1) Formulas may not always contain hydroxide ions
   iii) Strong acids/bases: strong electrolytes; weak acids/bases: weak electrolytes
      (1) Memorize the formulas of common strong acids & bases

Hydrochloric acid	HCl
Hydrobromic acid	HBr
Hydriodic acid	HI
Chloric acid	$HClO_3$
Perchloric acid	$HClO_4$
Nitric acid	HNO_3
Sulfuric acid	H_2SO4
Group 1 metal hydroxides	LiOH, NaOH, KOH, RbOH, CsOH
Heavy Group 2 metal hydroxides	$Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$

d) Neutralization Reactions & Salts
   i) Operational Definitions
      (1) Acids: sour taste, turn litmus paper red, react with metals to release H_2(g),
      (2) Bases: bitter taste, turn litmus paper blue, turn phenolphthalein pink
   ii) Neutralization reaction = when solution of an acid and a solution of a base are mixed
      (1) Acid + metal hydroxide → water + salt
      (2) Salt = any ionic compound whose cation comes from a base and whose anion comes from an acid = any ionic compound that is neither an acid nor a base
   iii) Acid-base reactions with gas formation
      (1) Sulfide ion → hydrogen sulfide gas (H_2S)
      (2) Carbonates, hydrogen carbonates → carbon dioxide (because carbonic acid is unstable and decomposes)
   iv) BE ABLE TO WRITE THESE AS NET IONIC EQUATIONS!

e) Oxidation-Reduction Reactions = reactions in which electrons are transferred between reactants = REDOX
   i) Oxidation = loss of electrons
   ii) Reduction = gain of electrons
   iii) Keep track of oxidation/reduction by using oxidation numbers   eActivity
      (1) The oxidation number of uncombined elements is zero.
      (2) For monatomic ions, the oxidation number equals the charge on the ion.
      (3) The oxidation number of oxygen is usually -2 in compounds; the major exception is the peroxide ion, which has an oxidation number of -1.
      (4) The oxidation number of hydrogen in compounds is +1, except in metal hydrides when it is -1.

(5) The oxidation number of fluorine is -1 in all compounds. The other halogens have an oxidation number of -1 in binary compounds but may have other oxidation numbers when oxygen is present.
(6) The sum of all the oxidation numbers in a neutral molecule is zero.
(7) The sum of all the oxidation numbers in a polyatomic ion is the charge of the polyatomic ion.

iv) Single Displacement Reactions: A + BX → AX + B Be able to write these as net ionic equations!
(1) Activity Series can be used to predict if a single displacement reaction will occur
(a) Active metals (at top of series) are easily oxidized; the "noble metals" at the bottom of the series have low reactivity
(b) A metal can replace any metal below it in the series in a single displacement reaction
(c) A similar series exists for halogen replacements
(d) You will not have (or need) an activity series for the AP Chemistry exam

f) Concentrations of Solutions = amount of solute dissolved in a given amount of solvent
i) Molarity $M = \dfrac{moles\ solute}{L\ of\ \ solution}$ critical calculation; be able to interconvert
ii) When ionic compounds dissolve, the relative concentration of the ions depends on the chemical formula (SUBSCRIPTS); so concentration of a particular ion in solution may equal or be greater than the solution concentration
iii) Dilution from a concentrated stock solution $M_{stock} V_{stock} = M_{dil} V_{dil}$ (Animation)
iv) Prove to yourself that moles solute taken from the stock solution = moles solute present in diluted solution! These are critical calculations; make sure you can do them reliably

g) Solution Stoichiometry & Chemical Analysis
i) These comparisons must be done on a mole: mole basis; avoid incorrect use of the dilution formula
ii) Titration = combining a sample of the solution with a reagent solution of known concentration (standard solution)
(1) Equivalence point = when stoichiometrically correct number of moles of each reagent is present
(2) End point = point where the indicator changes
(3) Indicator = weak acid or base that undergoes color change; choose indicator so that endpoint is close to equivalence point
(4) Acid-base, precipitation and redox titrations are common eAnimation, eActivity
(5) Problem solving is the key!

h) Oxidation States as "electronic bookkeeping" (jump to Ch. 20)
i) Oxidation = increase in oxidation number LEO the lion says GER
ii) Reduction = decrease in oxidation number OIL RIG

(1) Oxidation states of atoms in compounds do not necessarily show the actual charge of the atom in the compound (It's more along the lines of…"If this were an ionic compound, what would the ion charges be?")
(2) The substance that is oxidized is also called the reducing agent (or reductant)
(3) The substance that is reduced is also called the oxidizing agent (or oxidant)

b) Balancing Oxidation-Reduction Equations
  i) Half-reactions: showing oxidation and reduction alone, but we know that they always occur together!
  ii) To balance redox reactions, the number of electrons lost = the number of electrons gained
  iii) Balancing in acidic conditions:
    (1) Use oxidation numbers to identify the substances oxidized and reduced.
    (2) Write an incomplete, unbalanced half-reaction for each process.
    (3) Balance the elements other than H and O.
    (4) Balance the O atoms by adding H_2O as needed.
    (5) Balance the H atoms by adding H^+ as needed.
    (6) Balance the charge by adding electrons as needed.
    (7) Multiply the half-reactions by integers so that electrons lost = electrons gained.
    (8) Add the two half reactions and simplify.
  iv) Balancing in basic conditions:
    (1) Start as above, then add OH^- to neutralize the H^+ ions. You will make H_2O molecules, then simplify.