Things to Know for AP SkyviewChem

Things to Memorize

The following should be memorized. Not like you memorized the capital of Michigan or Turkey but the way you've memorized your birthday. Instantaneous recall is the standard you want to achieve. Don't dwell on anyone's opinion about the value of memorization, just do it. These things will not be the answers to AP questions but will lead you to the answers.

1. Ions – See the list below with the ions to memorize.
2. Solubility Rules – These are for ionic compounds and are not just about dissolving but about dissociating into ions. Chemically the difference in these two is shown by sugar and salt. $C_6H_{12}O_{6(s)} \rightarrow C_6H_{12}O_{6(aq)}$ while $NaCl_{(s)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$.
3. Common acids and bases – know which acids and bases are strong and which are considered weak
4. Assigning Oxidation States – This is the **imaginary** charge an atom has in a compound.
5. Oxidation and Reduction – know your strong oxidizing agents and strong reducing agents and the products which form from these reactions
6. Intermolecular Forces – know the different types and strengths
7. Organic Chemistry – know how to name simple organic compounds (ie ethane) and the main functional groups
8. Colors – know the colors of certain elements when they are burned (flame test) and when they are dissolved as ions. We will see some of these colors throughout the year in lab, and you can look them up on line.
9. Selected Elements and their Symbols – Periods 1-4.

Procedures to Practice

This short list of little skills is needed to solve problems. Usually they are things you do before the first step (identify the proper equation).

10. Calculate molar mass of an element or compound.
11. Problem Solving Rubric
12. Convert between grams, moles, Avogadro's number, and volume of a gas at STP.
13. Balance Chemical Equations.

1. Common Ions

Positive ions (cations)

+1 Charge

ammonium (NH_4^+)
copper (I) or cuprous (Cu^+)
hydrogen (H^+) "proton"
hydronium ion (H_3O^+)
silver (Ag^+)
Group 1 (Li^+, Na^+, K^+, Rb^+, Cs^+, Fr^+)

Negative ions (anions)

-1 Charge

acetate ($C_2H_3O_2^-$)
cyanide (CN^{-1})
dihydrogen phosphate ($H_2PO_4^-$)
hydrogen carbonate or bicarbonate(HCO_3^-)
hydrogen sulfate or bisulfate (HSO_4^-)
hydroxide (OH^-)
nitrate (NO_3^-), nitrite (NO_2^-)
perchlorate (ClO_4^-), chlorate (ClO_3^-), chlorite (ClO_2^-), hypochlorite (ClO^-)
permanganate (MnO_4^-)
Group 17, Halogens (F^-, Cl^-, Br^-, I^-)
thiocyanate (SCN^-)

+2 Charge

cadmium (Cd^{2+})
chromium (II) or chromous (Cr^{2+})
cobalt(II) or cobaltous (Co^{2+})
copper(II) or cupric (Cu^{2+})
iron(II) or ferrous (Fe^{2+})
lead(II) or plumbous (Pb^{2+})
manganese(II) or manganous (Mn^{2+})
Mercury(I) or mercurous ($Hg2^{2+}$)
Mercury(II) or mercuric (Hg^{2+})
nickel (Ni^{2+})
tin(II) or stannous (Sn^{2+})
zinc (Zn^{2+})
Group 2, Alkaline Earth Metals
($Be^{2+}Mg^{2+}$,Ca^{2+},Sr^{2+},Ba^{2+},Ra^{2+})

-2 charge

carbonate (CO_3^{2-})
Chromate ($CrO_4{}^{2-}$)
dichromate($Cr_2O_7^{2-}$)
Hydrogen phosphate (HPO_4^{2-})
oxalate ($C_2O_4^{2-}$)
oxide (O^{2-})
peroxide (O_2^{2-})
sulfate (SO_4^{2-})
sulfite (SO_3^{2-})
sulfide (S^{2-})
thiosulfate ($S_2O_3{}^{2-}$)

+3 Charge

aluminum (Al^{3+})
chromium(III) or chromic (Cr^{3+})
iron(III) or ferric (Fe^{3+})

-3 Charge

Arsenate (AsO_4^{3-})
phosphate (PO_4^{3-})
Phosphite (PO_3^{3-})
Group 15, pnictogen - nitride (N^{3-}), phosphide (P^{3-})

+4 Charge

lead(IV) or plumbic (Pb^{4+})
tin(IV) or stannic (Sn^{4+})

Summary of metal cations with more than one possible charge:

Cu^{+}, Cu^{2+}; Hg_2^{2+}, Hg^{2+}; Co^{2+}, Co^{3+}, Cr^{2+}, Cr^{3+}; Fe^{2+}, Fe^{3+}; Mn^{2+}, Mn^{3+}; Pb^{2+}, Pb^{4+}; Sn^{2+}, Sn^{4+}

Manganese and some other metals can form several ions with different charges. Know the ones listed.

2. Solubility Rules

1. All common compounds of Group 1 and ammonium ions are soluble.
2. All nitrates, acetates, and chlorates are soluble.
3. All binary compounds of the halogens (other than F) are soluble, **except** those of Ag, Hg(I), and Pb.
4. All sulfates are soluble, **except** those of barium, strontium, calcium, lead.
5. Sulfides and hydroxide are insoluble **except** for Ca, Ba, Sr, ammonium and the alkali metals.
6. Except for rule 1, carbonates, oxides, silicates, and phosphates are insoluble.

Note: You can apply the solubility rules to predict whether the product of a double replacement reaction will be a precipitate or not. If the compound is soluble, it will dissociate into free ions in solution. If the compound is insoluble, it will be a precipitate. Example: When $AgNO_3(aq)$ and $NaCl(aq)$ are mixed, the products will be AgCl and $NaNO_3$. AgCl is insoluble (rule 3) and will precipitate out. $NaNO_3$ is soluble and will remain in solution (rules 1 &2).

3. Acids and Bases

Common Acids

sulfuric acid - $H_2SO_4(aq)$

phosphoric acid – $H_3PO_4(aq)$

Carbonic acid – $H_2CO_3(aq)$

Sulfuric Acid - $H_2SO_4(aq)$

Acetic acid (ethanoic acid) $HC_2H_3O_2(aq)$ or $CH_3COOH(aq)$

7 Strong Acids

hydrochloric acid - $HCl(aq)$

hydrobromic acid – $HBr(aq)$

Hydroiodic acid - $HI(aq)$

nitric acid – $HNO_3(aq)$

Chloric acid - $HClO_3(aq)$

Chlorous acid - $HClO_4(aq)$

Common Bases

aqueous ammonia – $NH_3(aq)$

$RbOH(aq)$

7 Strong Bases

1A metal hydroxides (NaOH, LiOH…)

$Ca(OH)_2(aq)$

$Sr(OH)_2(aq)$

$Ba(OH)_2(aq)$

Note: Strong acids and bases are those that dissociate completely in water.

Example: $HCl(aq) \rightarrow {}^{+}(aq) + Cl^{-}(aq)$ $\quad$ $Ca(OH)_2(aq) \rightarrow Ca^{2+}(aq) + 2OH^{-}(aq)$

Weak acids and bases do not dissociate completely, and will be present as the compound in water.

Example: $HC_2H_3O_2(aq) \rightarrow H^{+}(aq) + C_2H_3O_2^{-}(aq)$ $\quad$ $NH_3(aq) \rightarrow NH_4^{+}(aq) + OH^{-}(aq)$

4. Rules for Assigning Oxidation Numbers

1. The oxidation number of any element in its free state (uncombined with other elements) is zero. e.g.,

Fe in a bar of iron is zero. O_2 and N_2 in the Earth's atmosphere both have oxidation numbers of zero. When an element has equal numbers of protons and electrons, its overall charge is zero.

2. The oxidation number of alkali metals **in a compound** is always 1+, e.g., LiCl, Na_3P, K_2S, etc.
3. The oxidation number of alkaline **earth** metals **in a compound** is always 2+, e.g., $MgCl_2$, CaF_2, SrO, etc.
4. Fluorine is always assigned a value of 1– **in a compound**, e.g., NaF
5. The oxidation number of oxygen is almost always 2– in a compound. Exceptions to this rule would be peroxides, O_2^{2-} where the oxidation number of **each** oxygen is 1–, and super peroxides, O_2^- where the oxidation number of each oxygen is ½–. Neither peroxides nor superperoxides are common. Peroxides are only known to form compounds with the elements in the first two columns of the periodic table, e.g., H_2O_2, Na_2O_2, CaO_2, etc. Potassium, rubidium, and cesium are the only elements that form superperoxides, e.g., KO_2.
6. In covalent compounds (with nonmetals), hydrogen is assigned an oxidation number of 1+, e.g., HCl, H_2O, NH_3, CH_4. The exception to this rule is when hydrogen combines with a metal to form a hydride. Under these conditions, which are rare, hydrogen is assigned an oxidation number of 1–, e.g., NaH.
7. In metallic halides the halogen (F, Cl, Br, I,) always has an oxidation number equal to 1–.
8. Sulfide, selenide, telluride are always 2– in binary salts.
9. Nitrides, phosphides, and arsenides are always 3– in binary salts.
10. All other oxidation numbers are assigned so that the sum of the oxidation numbers of each element equals the net charge on the molecule or polyatomic ion. In neutral compounds, the sum of the positive and negative charges must equal zero.

5. Oxidation and Reduction

Loss of electrons, Gain is reduction. (LEO goes GRRRRRR!!!!)

Strong Oxidizing agents Strong Reducing Agents

MnO_4^- in acid solution → Mn^{2+} + H_2O

MnO_4^- in neutral or basic solution → MnO_2

$Cr_2O_7^{2-}$ in acid solution → Cr^{3+} + H_2O

$Cr_2O_7^{2-}$ in basic solution → CrO_4^{2-} + H_2O

H_2SO_4 (concentrated) → SO_2 + H_2O

Free halogens (Cl_2) → Halide ions (Cl^-)

Halide ions (Cl^-) → Free halogens (Cl_2)

Free metals (Mg) → Metal ions (Mg^{2+})

"–ite" ions (SO_3^{2-}) → "–ate" ions (SO_4^{2-})

HNO_3 (concentrated) → NO_2 + H_2O

H_2O_2 in acid solution → H_2O

Study Suggestion: You may not completely understand this right now, but it will make sense later in the year. For now, learn the oxidizing and reducing agents in bold and the products formed when they react. Predicting products is very important in AP chem!

6. Intermolecular Forces

Forces between molecules that keep solids together.

Network Covalent Directional covalent bonds

C (graphite, diamond) Si, SiO_2(sand)

Strongest

Ionic (electrostatic attraction) Forces between adjacent ions ($Na^+ \rightarrow Cl^-$)

Metallic Forces between metal nuclei (Cu, Ag)

Hydrogen bonding Forces between adjacent molecules with H & F, O, N or Cl. (H_2O, NH_3)

Dipole-dipole Forces between adjacent polar molecules (CO, NF_3)

London Dispersion Force Forces between adjacent nonpolar molecules (CO_2, Cl_2)

Weakest

7. Organic Chemistry

Organic chemistry is the study of carbon based compounds. The nomenclature of organic compounds is based on the number of carbon atoms present in the molecule (prefix) and the presence of functional groups (ending).

examples) CH_4 – one carbon atom "meth-" & C_nH_{2n+2} … "methane"

C_2H_5OH - two carbon atoms "eth-" & an alcohol (-OH) … "ethanol"

$C_3H_7NH_2$ - three carbon atoms "prop-" & an amine (-NH_2) … "propylamine"

Functional Groups:

-ol

-al

-one

-oic acid

Benzene

C_6H_6

Prefix Number of C

meth- 1

eth- 2

prop- 3

but- 4

pent- 5

hex- 6

hept- 7

oct- 8

non- 9

dec- 10

C_nH_{2n+2}-ane C_nH_{2n}-ene C_nH_{2n-2}-yne

8. Colors:

Flame Test Colors Colors of some aqueous ions

Barium – green Co^{2+} pink

Calcium – orange Cu^{2+} blue green

Copper – blue/green Fe^{2+} olive green

Lithium – red Ni^{2+} bright green

Potassium – lavender Fe^{3+} brown

Sodium – yellow/orange CrO_4^{2-} orange

Strontium – orange/red $Cr_2O_7^{2-}$ yellow

10. Calculating the Molar Mass

Chemistry happens in moles and the easiest way to solve problems is using moles. Since we can't directly measure the number of moles we need to take something we can measure, mass, and turn that into moles. This makes calculating molar mass (the mass in grams of a substance which contains one mole of that substance) an important first step in many problems.

All year long we will use the official AP Chemistry Periodic Table for this and will report our molar mass to the hundredths place (X.XX). You will be given one and a copy of the official Periodic. You can print another if needed from the College Board website for AP Chemistry.

.

Use the complete atomic mass from the periodic table. Only round the final answer to the hundredths place.

Example: $Ca(NO_3)_2$

$Ca(NO_3)_2$ is 1 x Ca + 2 x N + 6 x O = molar mass

1 x 40.08 + 2 x 14.007 + 6 x 16.00 = 164.094 which rounds to 164.09 g/mol of $Ca(NO_3)_2$

11. Problem Solving Rubric for AP Chemistry

Many times during the year we will use equations to solve problems. You are given an equation sheet for the AP Chemistry Exam. Not all equations are on it so you will need to know some by heart. The grade you earn on problems which use equations is based on the following rubric.

- First point is for writing the proper equation.
- Second point is for rearranging the equation to solve for the unknown variable.
- Third point is for substituting the known values into the rearranged equation.
- Fourth point is for the answer with proper units & significant digits.

Example: What is the volume of 0.425 moles of helium at 1.10 atmospheres when the temperature is 20.5°C?

PV = nRT 1 point

V = (nRT)/P 1 point

V = (0.425)(0.0821)(293.5)/1.10 1 point

V = 9.31L 1 point

Longer problems and multipart problems will usually have some part or parts which use this rubric and will have more points for the parts before and after.