Teaching High School Chemistry – Core Concept Master Cheat Sheet

01: Introduction to High School Chemistry

What content are teachers covering?

571 high school teachers were surveyed on what chemistry topics were appropriate to teach. 96% of the teachers ranked these topics as appropriate.

1. Basic laboratory skills.
2. Basic skills
3. Dimensional analysis
4. Classification of matter.
5. Writing and naming formulas.
6. Moles
7. Types of reactions.
8. Balancing equations
9. Stoichiometry
10. Atomic structure (electron configuration)
11. Periodic table & periodicity
12. Types of bonds and properties.
13. Gas laws
14. Solutions and concentrations

Applying content to students' lives

Application of content increases motivation and interest, which in turn increases performance.

Textbooks that have true integration of application and introduce topics on a need-to-know basis.

- Chemistry In the Community (ChemCom) http://www.whfreeman.com/chemcom/
- Living by Chemistry http://www.keypress.com/chemistry/
- Active Chemistry http://www.its-about-time.com/htmls/ac/ac.html
- Chemistry in Your World
- www.RealLifeChemistry.net

Ideas for research projects and presentations applying chemistry to students' lives:

- Fuel-cell cars
- Hair dyes
- Fireworks
- Luminol
- Glow in the dark algae.
- Nuclear power plant.
- Or anything else imaginable.

Articles with application:

- Newspapers
- Chemical & Engineering News
- ChemMatters

02: Teaching Labs in Chemistry

Lab Safety

- Wear splash-proof goggles.
- No eating or drinking.
- Do not touch, taste or directly smell chemicals.
- Tie back all loose clothing, hair and jewelry.
- Always read the procedure ahead of time and follow it closely.
- Never return unused chemicals to the original container.
- Dispose of all chemicals as instructed.
- Always report all incidents (spills, breakage, mistakes in performing a procedure) to your instructor!

Running a Lab

- Always include written safety information and go over it verbally.
- Go over any new techniques or equipment ahead of time
- Walk around and keep an eye on everything!
- Refill chemical supply when needed—but not too much to keep waste down.
- Ask students questions to engage them with the procedure.
- Anticipate questions—if you see one group having a question or problem, others will as well!

03: Assessments in Chemistry

- **Long answer:** Assesses critical thinking and application of knowledge. Essay, short answer essay and calculations problems.
- **Short answer:** Assess knowledge of facts. True/false, matching, fill in the blank and multiple choice.

Ways to increase difficulty of short answer

- **Multiple-choice:** Require correct work shown on quantitative questions to eliminate awarding points for correct guessing.
- **Matching:** Have more options than questions so they cannot apply "process of elimination" on the last few.
- **True/False:** Have them correct false statements, or explain why they're false
- **Fill in the bank:** Don't provide a word bank.

04: Measurement & Math in Chemistry

The **metric system** uses prefixes to indicate multiples of 10.

Metric Prefixes commonly used in chemistry

Prefix	Symbol	Multiple
Kilo	k	1000
Deci	d	0.1
Centi	c	0.01
Milli	m	0.001
Micro	μ	0.000001
Nano	n	0.000000001

The "base unit" is when there's no prefix.

The **SI sytem** gives the fundamental unit for each type of measurement.

Counting Significant Figures:

- If there is a decimal point anywhere in the number: Start with the first non-zero number and count all digits until the end.
- If there is not a decimal point in the number: Start with the first non-zero number and count until the last non-zero number.

Calculations with significant figures:

- Always complete calculations before rounding.
- Adding/subtracting: Answer has least number of decimal places as the problem.
- Multiplying/dividing: Answer has least number of significant figures in problem.

Scientific Notation—a short hand method of writing numbers using powers of 10.

Writing scientific notation:

1. The decimal point is always moved to after the 1st non-zero number.
2. Count the number of times the decimal point is moved and use this as the power of 10.
3. "Big" numbers (>1) have positive exponents. "Small" numbers (<1) have negative exponents.

Reading scientific notation:

1. Power of 10 = number of times to move decimal point
2. Positive powers = make the number "Big" (>1). Negative exponents = make the number "Small" (<1).

Logarithms: Way of counting in multiples of the base

$$x = \log_b y$$

Calculator tips:

- Always use the ÷ key to designate a number is on the bottom of an expression.
- Always use the EE (or EXP) key to enter scientific notation.
- Always use parenthesis around addition or subtraction when combining it with other operations.
- To make something negative (when taking the number to a power), keep the negative outside of the parenthesis.

05: Dimensional Analysis

Dimensional analysis is the technique used to convert units.

The principle behind dimensional analysis:
Multiplying by 1 does not change the physical meaning of the measurement.

Using Dimensional Analysis:
1. Write your given information on the left side.
2. Write "= ______ (desired unit)" on the right side.
3. Find equalities that include both the desired unit and the given unit.
4. Arrange the equalities so that the given unit cancels.
5. Calculate answer, multiply across the top and divide across bottom.

Multi-step Dimensional Analysis
- If there is no equality that contains both the given and the desired unit, you will need to use more than one equality.
- If you convert from a metric prefix to another metric prefix, use the base unit as a bridge in-between.

When converting a quantity with a fractional unit:
Separate the unit—put the top on the top of the expression and the bottom of the unit on the bottom of the expression.

06: Solving Chemistry Problems

Use the KUDOS method for solving word problems.
K = Known
U = Unknown
D = Definition
O = Output
S = Substantiation

- **K (Known):** Use units to indentify information, Write information symbolically, Look for implied information, Write out chemical equations.

- **U (Unknown):** What is the problem looking for? Write information symbolically.

- **D (Definition):** Find equalities to convert. Choose and rearrange equations. Look for missing information in other places. If you cannot find enough information, re-evaluate your plan.

- **O (Output):** Plug in values to the equations, use constants as needed. Check unit cancellation and perform the calculation.

- **S (Substantiation):** Check validity of your answer. Check units and significant figures.

O7: Energy & Matter

Matter			
Pure Substance		Mixtures	
Element Hydrogen	Compound H2O	Homogeneous Tap water	Heterogeneous Sand & Water

Energy	
Kinetic Energy (KE) Energy due to motion	Potential Energy (PE) Stored in chemical bonds

Physical Changes
- <u>Do not</u> create a new substance.
- All changes in state (between solids, liquids and gases) are physical changes.

Breaking, cutting, dissolving, drying, melting, freezing, etc.

Chemical Changes
- <u>Do</u> produce new substances.

Some signs of a chemical change are:
- Production of a gas (bubbles)
- Heat change, getting hot or cold.
- Light
- Change in color.
- Formation of a precipitate (forming an insoluble substance from two soluble substances.
- However, some of these signs could be present in physical changes as well.
- Rusting, burning, reacting with water, reacting with acid, etc.

08: Pure Substances—Atoms & Molecules

Sub-atomic Particles:

Particle	Location	Mass	Charge
Proton	Nucleus	1 amu = 1.67×10^{-27} kg	+1
Neutron	Nucleus	1 amu = 1.67×10^{-27} kg	0
Electron	Outside the nucleus	0.00055 amu = 9.10×10^{-31} kg	-1

Ions
- Atoms can gain or lose electrons to form **ions** (atoms with a charge.
- **Anion:** Atom with a negative charge.
- **Cation:** Atom with a positive charge.

Element symbols:

$${}^{A}_{Z}X^{C}_{\#} \quad \text{Where}$$

- A = mass number (# of protons + # of neutrons)
- Z = atomic number (# of protons)
- C = charge (# of protons - # of electrons)
- # = number of atoms

Isotopes: Atoms of same element with different number of neutrons (and different mass).
- Mass number refers only to a specific isotope.

Calculating average atomic mass: Found on periodic table.
Atomic mass = Σ(fractional abundance)(mass of that isotope).

Atoms, elements and molecules
- **Atoms:** Made of sub-atomic particles.
- **Elements:** Made of the same type of atom (each has the same number of protons).
- **Molecules:** Made of more than one type of atom (more than one element) chemically bonded together.

09: Writing Chemical Formulas

Type 1 Binary ionic: Contains two elements—one metal and one non-metal.

1. Write the symbol and charge of the first element.
2. Write the symbol and charge of the second element.
3. Balance the charges to form a neutral compound by using subscripts.

Type 1 or 2 with Multivalent Metals

Metals that can have more than one charge.

1. The Roman numeral indicates the charge of the cation metal.
2. Follow the rules for Type #1 or Type #2 as it applies.

Type 3 Binary Covalent: Contains two non-metals (which do not form charges when bonding together).

1. Do not worry about charges with this type.
2. Write the first element's symbol.
3. Write the second element's symbol.
4. Use the prefixes to determine subscripts ("mono" is not used on the first element).

Acids:

1. "Acid" indicates "H^+" is the cation.
2. Choose the anion:
   a. "hydro__ic acid" – anion is single element (no oxygen).
   b. "__ic acid" – anion is "__ate" ion
   c. "__ous acid" – anion is "__ite" ion
3. Balance charges with subscripts.

10: Naming Chemicals

Type 1 Binary ionic: Contains two elements—one metal & one non-metal.

1. Write the name of the first element.
2. Write the name of the second element with "-ide"

(subscripts do not matter in this type).

Type 2 Polyatomic Ionic: Contains at least one **polyatomic ion** (group of atoms that together have a charge).

1. Write the name of the metal or "ammonium" for NH_4.
2. Write the name of the polyatomic anion (do not change the ending) or the single element with "-ide".

Subscripts within a polyatomic ion must match the name exactly. If there are parenthesis, the polyatomic ion is inside the parenthesis.

Type 1 or 2 with Multivalent Metals

Metals that can have more than one charge.

Co, Cr, Cu, Fe, Hg, Pb, Sn

1. Name the cation and anion as for Type #1 or Type #2.
2. The compound is neutral. Use the charge of the anion to determine the charge of the cation.
3. Write the charge of the cation in Roman numerals inside parenthesis.

Type 3 Binary Covalent: Contains two non-metals (which do not form charges when bonding together).

1. Write the first element's name with the prefix indicating the # of molecules (mono- is not used with the first element).
2. Write the second element's name with the prefix indicating the # of molecules and "-ide".

Acids: (Compounds with "H^+" cations are acids)

1. Look up the anion:
   a. No oxygen, a single element: "hydro__ic acid"
   b. "__ate" ion: "__ic acid"
   c. "__ite" ion: "__ous acid"

10: Counting Molecules—The Mole

Mole: SI unit for counting (abbreviation: mol)

- 1 mole of anything = 6.02×10^{23} pieces.
- The atomic mass found on the periodic table is the mass (in grams) for 1 mole of atoms of that element.
- At standard temperature and pressure (STP), 1 mole of any gas is 22.4 L **(Molar Volume of a gas)**

Molar Mass (Molecular Mass, Formula Weight):

- By adding the atom masses for atoms in a molecule, the molar mass of the molecule can be found.
- Be sure to distribute subscripts outside the parenthesis to each atom inside.

Percent Composition:

$$\%\ composition = \frac{mass\ element}{mass\ whole} \times 100$$

If a chemical formula is given, use atomic masses and molar mass in % composition.

Empirical formula (lowest ratio of atoms in molecule):

1. If given percent's, assume they are grams. Change all grams to moles.
2. Divide all moles by the smallest to get the lowest ratio. Multiply by a factor if needed to make whole numbers.
3. Write the formula with the ratio as subscripts.

Molecular Formula (actual ratio of atoms in a molecule):

1. Find empirical formula, if not given to you.
2. Find the molar mass of the empirical formula.
3. Find the ratio of the molecular formula's molar mass (given to you) to the empirical formula's molar mass.
4. Multiple the empirical formula's subscripts by the ratio.

12: Chemical Reactions

Chemical Reaction: Bonds and atoms are rearranged to form new compounds.

Chemical Equation: Symbolizes the chemical reaction with chemical formulas.

- Reactants → Products
- States of matter are shown (s = solid, l = liquid, g = gas, aq = aqueous).
- Coefficients give mole ratio.
- A double arrow (⇆) indicates it is a reversible reaction also known as an equilibrium reaction.

Types of reactions:

- **Composition:** More than one type of matter combine to form one type of matter.
- **Decomposition:** One type of matter decomposes into more than one type of matter.
- **Single replacement:** A single element changes place with an ion in a compound.
- **Double replacement:** Two ionic compounds switch ions.
- **Neutralization reaction:** Double replacement reaction with an acid and a base as the reactants.
- **Redox reaction:** Reduction-oxidation reaction.
- **Precipitation reaction:** A precipitate is formed.

Solubility rules for determining precipitates:

Anion	Forms insoluble compounds with
NO_3^-	No common ions
CH_3COO^-	Ag^+
Cl^-, Br^-, I^-	Ag^+, Pb^{2+}, Hg_2^{2+}, Ti^+
SO_4^{2-}	Ag^+, Pb^{2+}, Ba^{2+}, Sr^{2+}, Ca^{2+}
CrO_4^{2-}	Ag^+, Pb^{2+}, Ba^{2+}, Sr^{2+}
S^{2-}	All anions except NH4+, columns 1 & 2
OH^-	All anions except NH_4^+, column 1, Ba^{2+} & Sr^{2+}
CO_3^{2-}, PO_4^{3-}	All anions except NH_4^+, column 1 (except Li^+)

NH_4^+, Na^+ and K^+ are soluble with all common ions

13: Balancing Equations

- The Law of Conservation of Mass/Matter requires that a chemical reaction be balanced.
- Coefficients balance atoms in a chemical reaction and indicate the number of compounds in a reaction.

Inspection Method (to balance the most simple reactions):

1. Make a list of the elements in the reaction.
2. Count the # of each type of atom on each side.
3. Add coefficients to balance the number of atoms.
4. Determine the total charge of each side and use coefficients to balance charge.
5. When elements and charge are balanced, place a “1” in any empty coefficient location.

$Zn + H^+ \rightarrow Zn^{2+} + H_2$

	Reactants	Products
Zn	1	1
H	1	2
Charge	+1	+2

__ Zn + _2_ $H^+ \rightarrow$ __ Zn^{2+} + __ H_2

	Reactants	Products
Zn	1	1
H	~~1~~ 2	2
Charge	~~+1~~ 2	+2

$1\ Zn + 2\ H^+ \rightarrow 1\ Zn^{2+} + 1\ H_2$

14: Stoichiometry

Stoichiometry: Using the mole ratio in the balanced equation and information about one compound to find information about another in the reaction.

Equalities used during dimensional analysis for stoichiometry:

- **Mole ratio in balanced equation:** Used to convert between moles of different compounds in the balanced equation.
- **Molar mass:** Used to convert between grams and moles.
- **Concentration:** Used to convert between moles and liters of a solution.

$$Molarity = \frac{moles\ solute}{L\ solution}$$

- **Molar volume of a gas:** Used to convert between moles and liters of a gas at STP.

Mass-Mass example:

If 2.5 g Mg react, how many grams $MgCl_2$ are produced?

2.5 ~~g Mg~~	1 ~~mole Mg~~	1 ~~mole $MgCl_2$~~	95.21 g $MgCl_2$	= 9.8 g $MgCl_2$
	24.31 ~~g Mg~~	1 ~~mole Mg~~	1 ~~mole $MgCl_2$~~	

Limiting reactant: Reactant that stops the reaction by running out first.

- Once a reactant has run out, the reaction will stop.
- Do stoichiometry for each given reactant quantity to the same product each time. Choose the calculation that gives the smallest amount of product.
- The reactant that produces the smallest amount of product is the limiting reactant.

Percent yield: compares the actual yield to the theoretical yield.

$$\%\ yield = \frac{actual\ yield}{theoretical\ yield} \times 100$$

15: Electron Configuration

Electron cloud: Area outside nucleus where electrons are located.
Energy levels: Electron cloud is divided into energy levels for electrons.
Subshells: Energy levels of electrons are divided into subshells of equal energy orbitals.
Orbitals: Subdivision of a subshell. Each orbital can hold 2 electrons.

4 types of subshells:

	Subshell	Begins in level	# of orbitals	# of electrons
↓ higher energy	s	1	1	2
	p	2	3	6
	d	3	5	10
	f	4	7	14

Aufbau Principle: Fill shells from lowest energy to highest.
Hund’s Rule: Electrons are placed in each equal-energy orbital before doubling up to produce the lowest energy atom.
Pauli Exclusion Principle: Two electrons occupying the same orbital must have opposite spins (angular momentum).

Use the periodic table as a guide (read left to right):

1s			
2s			2p
3s			3p
4s		3d	4p
5s		4d	5p
6s	4f	5d	6p
7s	5f	6d	7p

3 types of electron configuration notation:

Boxes & Arrows: O (8 electrons): 1s $\boxed{\uparrow\downarrow}$ 2s $\boxed{\uparrow\downarrow}$ 2p $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}$ $\boxed{\uparrow}$
Spectroscopic: Br (35 electrons):
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$
Noble Gas: Br (35 electrons): $[Ar]\ 4s^2\ 3d^{10}\ 4p^5$

16: The Periodic Table

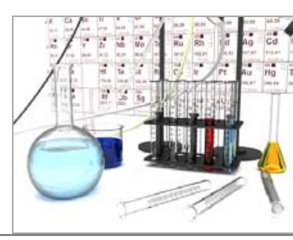
Periodic Table: Tool for organizing the elements.
Periods: Rows on the periodic table.
Groups: Columns on the periodic table.
Periodicity: Predictable patterns and trends on the periodic table.

General trends in the period table

Trend	→ a period	↓ a group
Atomic Mass	Increases	Increases
Atomic Radii	Decreases	Increases
Ionization energy	Increases	Decreases
Electron Affinity	Increases	Decreases
Electronegativity	Increases	Decreases

Radii when forming a cation: There are now more protons than electrons. The pull of the protons on each electron is greater. Cations have smaller radii than their parent atom.

Radii when forming an anion: There are fewer protons than electrons. The pull of the protons on each electron is less.
Anions have larger radii than their parent atom.



17: Chemical Bonding

Bond type	Happens between	Electrons are
Ionic	Metal & non-metal	Transferred
Covalent	Non-metals	Shared
Polar Covalent	Non-metals	Shared unevenly
Metallic	Metals	pooled

Polar covalent bond
When nonmetals bond covalently with a large difference in electronegativity.

- Absolute value of differences:
    d. 0 – 0.4 = covalent
    e. 0.5 – 1.4 = polar covalent
    f. 1.5 – 4 = ionic

Sigma (σ) bond: First bond between two atoms formed from head on overlap of orbitals
Pi (π) bond: bond between two atoms formed from overlap of parallel p orbitals.

Each single bond is a sigma bond.

Each double or triple bond contains one sigma bond and then pi bonds to form the second or third bond.

Example:
How many sigma and pi bonds are in the following?

```
     H   H
     |   |
H – C = C – C ≡ C
```

6 sigma bonds & 3 pi bonds

18: Molecular Structures

Valence Shell: Electrons in the outermost shell that bond.
Octet Rule: Atoms are most stable when having a full valence shell.
Arranging Atoms in Lewis Structures

1. With only 2 elements, arrange symmetrically.
2. "COOH" is a carboxylic acid. Both O's bond to the C and the H goes on one of the O's.
3. Hydrogen and halogens cannot go in the middle.
4. Other atoms in the order they appear in the formula
5. Hydrogen and halogen atoms go around the element they are written next to in the formula.

Lewis Structure: A 2D representation of a molecule and its bonds.

1. Arrange the atoms as above.
2. Determine the # of valence electrons for each atom.
3. Draw the valence electrons—do not double up where a bond is going to form between two atoms.
4. Count to see if all atoms have full valences.
5. If two atoms adjacent to each other do not have full valences, move in an electron from each to form a double bond. Repeat for triple bond if necessary. Move hydrogens as needed to allow double/triple bonds.

Exceptions to the Octet Rule:

1. Hydrogen and Helium can only hold 2 electrons Boron and Beryllium can be full with 6 electrons.
2. Any element in period 3 or below can have more than 8 electrons.

Valence Shell Electron Pair Repulsion Theory (VSEPR):
Bonds and lone pairs (electrons) repel and arrange themselves in 3D as far away from each other as possible.

19: Gas Laws

Assumptions of the KMT

1. Gases are made of atoms or molecules.
2. Gas particles are in rapid, random, constant motion.
3. The temperature is proportional to the average kinetic energy.
4. Gas particles are not attracted nor repelled from each other
5. All gas particle collisions are perfectly elastic (they leave with the same energy they collided with).
6. The volume of gas particles is so small compared to the space between them, that the volume of the particle is insignificant.

Symbols for all gas Laws:
P = Pressure; V = Volume; n = moles; T = Temperature (in Kelvin: K = °C + 273); R = Gas constant (8.31 L kPa/mole K or 0.0821 L atm/mole K); "a" and "b" = correction factors for real gases.

Combined Gas Law: $\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$

Dalton's Law of Partial Pressure: $P_{total} = \sum P_{of\ each\ gas}$

Mole fraction: $\chi_A = \frac{mole_A}{mole_{total}}$

Partial Pressure and mole fraction: $P_A = \chi_A P_{total}$

Ideal Gas Law: $PV = nRT$

Real Gas Law: $\left(P + \frac{n^2 a}{V^2}\right)(V - nb) = nRT$

20: Solutions

Solution: Homogeneous mixture.

Solute: Substance being dissolved.

Solvent: Substance doing the dissolving.

Factors affecting Solubility:

- **Pressure:** Gases: as Pressure increases, solubility increases
- **Temperature:** Gases: higher temperature is lower solubility. Most solids: higher temperature is higher solubility.

Concentration Measurements:

% by mass: $\%\ mass = \frac{mass\ solute}{mass\ solution} \times 100$

Molarity (M): $Molarity = \frac{moles\ solute}{L\ solution}$

Molality (m): $Molality = \frac{moles\ solute}{kg\ solvent}$

Dilution equation: $M_1V_1 = M_2V_2$

Electrolyte: Compounds dissociate into ions when dissolved in water. This allows the solution to conduct electricity.

21: Reaction Rates & Equilibrium

Kinetics: The study of reaction rates.

In order for a reaction to occur, the molecules must:
Collide with the correct orientation and activation energy.
Activation Energy is the minimum energy needed for a reaction to occur.

Factors affecting rate:
- **Surface area**—As surface area increases, rate increases.
- **Concentration**—As concentration increases, rate increases.
- **Temperature**—As temperature increases, rate increases.
- **Catalyst**—Presence of a catalyst increases rate.

Reversible Reaction: Reaction that goes in both directions.

Equilibrium: When the rate of the forward and reverse of a reversible process are equal.

Dynamic equilibrium: The number of reactants and products do not change, but the reaction continues to occur in both directions.

Writing Equilibrium Constant Expressions
- Concentration of products over concentration of reactants.
- Do not include pure solids or pure liquids.
- Use the coefficients of the balanced equations as powers.

Reaction Quotient (Q): When concentrations at any time are plugged into the equilibrium constant expression.
- If Q = K, it's at equilibrium.
- If Q > K, reaction proceeds towards reactants.
- If Q < K, reaction proceeds towards products.

Le Chatelier's Principle: A system at equilibrium will re-adjust to reach equilibrium again when disturbed.

22: Acids and Bases

Arrhenius acid: Produces hydronium ion in water.

Arrehnius base: Produces hydroxide ion in water.

Hydronium ion: H_3O^{+1}; **Hydroxide ion:** OH^{-1}

Strong acids	HCl, HBr, HI, HNO_3, $HClO_3$, $HClO_4$
Strong bases	$NaOH$, KOH, $Ca(OH)_2$, $Ba(OH)_2$, $Sr(OH)_2$

Polyprotic acids: each successive proton is weaker than the one before. (H_2SO_4 has a strong 1^{st} hydrogen).

pH: Logarithmic scale of acidity.

$$pH = -\log[H_3O^{+1}]$$

$$[H_3O^{+1}] = 10^{-pH}$$

$$K_w = [H_3O^{+1}][OH^{-1}]$$ at 25°C, $K_w = 1.0\times10^{-14}$

Salt from
- Weak acid + strong base = Basic
- Strong acid + weak base = Acidic

Buffer: Weak acid or base and its conjugate that resists changes in pH when acid or bases is added.

23: Thermodynamics

Thermodynamics: Study of heat changes.
Energy: The ability to do work or supply heat.
Heat (q): Flow of energy from a hotter object to a cooler object.
Enthalpy (H): Takes into account internal energy, pressure and volume. Same as heat for open-air situations.
Calorimetry: $\Delta H_{system} = \Delta H_{surroundings}$
T_2 of both system and surroundings are the same

Work $w = -P\Delta V$

w = work (in J); P = pressure (in atm); $\Delta V = V_2 - V_1$ (in L)

For changes in temperature:

$$\Delta H = m \times C_p \times \Delta T$$

m = mass
$\Delta T = T_2 - T_1$

For changes in state: Temperature doesn't change as the added energy is used to break intermolecular forces.

Melting: $\Delta H = m \times H_{fus}$

H_{fus} = enthalpy of fusion

Boiling: $\Delta H = m \times H_{vap}$

H_{vap} = enthalpy of vaporization
(freezing and condensing use the opposite values—exothermic)

Enthalpy of formation (H_f): Energy change when a compound is formed from its elements.

$$\Delta Hrxn = \sum H_f\, prod - \sum H_f\, react$$

Entropy (S): Disorder or random-ness
Free Energy (G): Takes into account enthalpy, entropy and temperature to determine spontaneity

$$\Delta G = \Delta H - T\Delta S$$

- ΔG = Spontaneous at that temperature.
+ ΔG = Spontaneous in the opposite direction at that temperature.

24: Electrochemistry

Electrochemistry: The study of the inter-change between electrical and chemical energy.
Voltaic cell (or Galvanic cell): Uses a redox reaction to produce electricity.
Electromotive force, EMF (or Cell Potential): Difference of potential energy of electrons from before and after the transfer.
Standard reduction potential: EMF if hydrogen is used as the other half-reaction (Hydrogen is defined as "0").
Calculating EMF from standard reduction potentials:
EMF = cathode – anode + EMF = spontaneous
Stoichiometry & Electrochemistry:
1 amp (A) = 1 Coulomb/sec (C/s)
1 Faraday (F) = 1 mole of e^{-1}
1 Faraday (1 mole of e^{-1}) = 96475 Coulomb (C)
Oxidation number rules:
- The sum of all oxidation numbers must equal the overall charge of the species. 0 for elements or compounds, the charge for a polyatomic ion.
- Hydrogen is +1 when with nonmetals, -1 with metals.
- Oxygen is usually -2.
- Halogens (column 7) are usually -1.
- The oxidation number of an ion in an ionic compound is the charge.

For redox reactions that cannot be balanced with inspection method:
- Determine the oxidation numbers of each atom.
- Determine the net change in charge. Use the net change to determine the ratio of atoms that would cancel out the net charge change.
- Use the ratio as coefficients in the simplest compounds containing those elements.
- Finish balancing by the inspection method.

