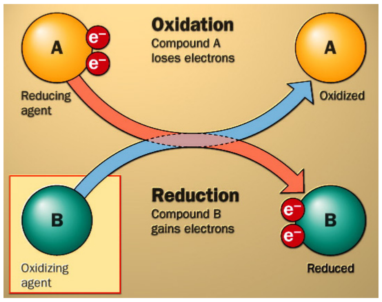


- **Oxidation** is the loss of electrons during a reaction by a molecule, atom or ion. Oxidation occurs when the oxidation state of a molecule, atom or ion is increased. The opposite process is called reduction, which occurs when there is a gain of electrons or the oxidation state of an atom, molecule, or ion decreases.

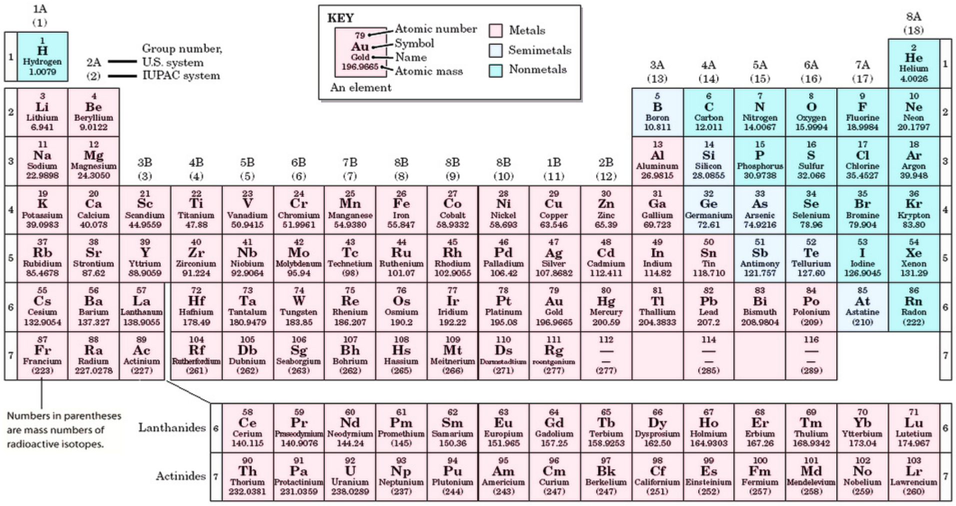
-The **valence** of an element is the measure of its combining capacity with other atoms when it forms chemical compounds or molecules.

-A half **reaction** is either the oxidation or reduction reaction component of a redox reaction. A half reaction is obtained by considering the change in oxidation states of individual substances involved in the redox reaction

-**Titration** is a process for ascertaining the quantity of any given constituent present in a compound by observing the quantity of a liquid of known strength (called a *standard solution*) necessary to convert the constituent into another form, the close of the reaction being marked by some definite phenomenon, usually a change of color or the formation of a precipitate. Also called volumetric analysis.

-**Periodicity**, the repetition of patterns at regular intervals, is the most important property of the Periodic Table. In the Periodic Table, pure elements are ordered in increasing order of atomic number, which is the number of protons in nucleus. Other important factors in determining the arrangement of the elements include electron configurations and recurring chemical and physical properties.

These **periodic trends** may increase, decrease, or stay consistent along a row or column. The periodic table also contains four rectangular blocks with approximately similar properties. **Metals** tend to be found more towards the left of the table while **nonmetals** tend to be found on the right, with **metalloids** in between, following a "staircase" pattern.



KEY
79 — Atomic number
Au — Symbol
Gold — Name
196.9665 — Atomic mass
An element

🟥 Metals
▢ Semimetals
🟦 Nonmetals

Group number,
2A — U.S. system
(2) — IUPAC system

	1A (1)	2A (2)	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	8B (8)	8B (9)	8B (10)	1B (11)	2B (12)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)	
1	1 **H** Hydrogen 1.0079																	2 **He** Helium 4.0026	1
2	3 **Li** Lithium 6.941	4 **Be** Beryllium 9.0122											5 **B** Boron 10.811	6 **C** Carbon 12.011	7 **N** Nitrogen 14.0067	8 **O** Oxygen 15.9994	9 **F** Fluorine 18.9984	10 **Ne** Neon 20.1797	2
3	11 **Na** Sodium 22.9898	12 **Mg** Magnesium 24.3050											13 **Al** Aluminum 26.9815	14 **Si** Silicon 28.0855	15 **P** Phosphorus 30.9738	16 **S** Sulfur 32.066	17 **Cl** Chlorine 35.4527	18 **Ar** Argon 39.948	3
4	19 **K** Potassium 39.0983	20 **Ca** Calcium 40.078	21 **Sc** Scandium 44.9559	22 **Ti** Titanium 47.88	23 **V** Vanadium 50.9415	24 **Cr** Chromium 51.9961	25 **Mn** Manganese 54.9380	26 **Fe** Iron 55.847	27 **Co** Cobalt 58.9332	28 **Ni** Nickel 58.693	29 **Cu** Copper 63.546	30 **Zn** Zinc 65.39	31 **Ga** Gallium 69.723	32 **Ge** Germanium 72.61	33 **As** Arsenic 74.9216	34 **Se** Selenium 78.96	35 **Br** Bromine 79.904	36 **Kr** Krypton 83.80	4
5	37 **Rb** Rubidium 85.4678	38 **Sr** Strontium 87.62	39 **Y** Yttrium 88.9059	40 **Zr** Zirconium 91.224	41 **Nb** Niobium 92.9064	42 **Mo** Molybdenum 95.94	43 **Tc** Technetium (98)	44 **Ru** Ruthenium 101.07	45 **Rh** Rhodium 102.9055	46 **Pd** Palladium 106.42	47 **Ag** Silver 107.8682	48 **Cd** Cadmium 112.411	49 **In** Indium 114.82	50 **Sn** Tin 118.710	51 **Sb** Antimony 121.757	52 **Te** Tellurium 127.60	53 **I** Iodine 126.9045	54 **Xe** Xenon 131.29	5
6	55 **Cs** Cesium 132.9054	56 **Ba** Barium 137.327	57 **La** Lanthanum 138.9055	72 **Hf** Hafnium 178.49	73 **Ta** Tantalum 180.9479	74 **W** Tungsten 183.85	75 **Re** Rhenium 186.207	76 **Os** Osmium 190.2	77 **Ir** Iridium 192.22	78 **Pt** Platinum 195.08	79 **Au** Gold 196.9665	80 **Hg** Mercury 200.59	81 **Tl** Thallium 204.3833	82 **Pb** Lead 207.2	83 **Bi** Bismuth 208.9804	84 **Po** Polonium (209)	85 **At** Astatine (210)	86 **Rn** Radon (222)	6
7	87 **Fr** Francium (223)	88 **Ra** Radium 227.0278	89 **Ac** Actinium (227)	104 **Rf** Rutherfordium (261)	105 **Db** Dubnium (262)	106 **Sg** Seaborgium (263)	107 **Bh** Bohrium (262)	108 **Hs** Hassium (265)	109 **Mt** Meitnerium (266)	110 **Ds** Darmstadtium (271)	111 **Rg** roentgenium (277)	112 **—** — (277)		114 **—** — (285)		116 **—** — (289)			7

Numbers in parentheses are mass numbers of radioactive isotopes.

Lanthanides	6	58 **Ce** Cerium 140.115	59 **Pr** Praseodymium 140.9076	60 **Nd** Neodymium 144.24	61 **Pm** Promethium (145)	62 **Sm** Samarium 150.36	63 **Eu** Europium 151.965	64 **Gd** Gadolium 157.25	65 **Tb** Terbium 158.9253	66 **Dy** Dysprosium 162.50	67 **Ho** Holmium 164.9303	68 **Er** Erbium 167.26	69 **Tm** Thulium 168.9342	70 **Yb** Ytterbium 173.04	71 **Lu** Lutetium 174.967	6
Actinides	7	90 **Th** Thorium 232.0381	91 **Pa** Protactinium 231.0359	92 **U** Uranium 238.0289	93 **Np** Neptunium (237)	94 **Pu** Plutonium (244)	95 **Am** Americium (243)	96 **Cm** Curium (247)	97 **Bk** Berkelium (247)	98 **Cf** Californium (251)	99 **Es** Einsteinium (252)	100 **Fm** Fermium (257)	101 **Md** Mendelevium (258)	102 **No** Nobelium (259)	103 **Lr** Lawrencium (260)	7

Group 1: Alkali Metals- Shiny, low densities, high conductivity, very low melting points, low electron affinities (Valence Electrons: 1)They are extremely light, reactive, and soft metals. Some spontaneously burn in air, and they all react violently on contact with water.

Group 2: Alkaline Earth Metals- Shiny, silvery-white, high conductivity, low melting points, low electron affinities (Valence Electrons: 2) **They** are slightly less reactive than the alkali metals, but are still reactive and flammable, but less so than their neighbors.

Group 3-12: Transitional Metals- Metallic, high densities, high conductivity, moderately reactive, high melting points, low electron affinities (Valence Electrons: 3-8) **They** are what most people think of when they hear the word "metal." They are generally tough, stable, and hard to melt, with some exceptions. Technetium (Tc, 43) is radioactive, along with all the group 7 transition metals, and mercury (Hg, 80) is liquid at room temperature.

Group 11: Coinage Metals- Metallic, nonreactive, high melting points (Valence Electrons: 4) They are a subgroup in the transition metals. They are very good conductors of electricity, and silver (Ag, 47) is the best conductor of any element.

Group 12: Volatile Metals- Metallic, high conductivity, moderately reactive, (Valence Electrons: 2, 4) They are a subgroup of the transition metals. They have lower melting points than the other transition metals, and its fourth member, the highly radioactive element opernicium, is even predicted to be a metal gas.

Group 13: Boron Group- Diverse properties, moderately reactive (Valence Electrons: 3) The properties vary by element, although all can react as +3. As you go down the column, the likelihood of reacting as +1 increases.

Group 14: Carbon Group- Diverse properties, moderately nonreactive (Valence Electrons: 4) Properties of the **carbon group** elements vary widely.

Groupe 15: Pnictogens (Nitrogen Family)- Diverse properties, moderately reactive (Valence Electrons: 5) The **pnictogens'** properties vary widely as well.

Group 16: Chalcogens (Oxygen Family)- Diverse properties, reactive, high electron affinities (Valence Electrons- 6). Name comes from Greek-"ore formers." Like the boron group, the **chalcogens** are generally moderately reactive, but have variable properties.

Group 17: Halogens- Highly reactive, very high electron affinities (Valence Electrons- 7)

Group 18: Noble Gases- Gaseous, odorless, colorless, inert (valence electrons- 0, 2, 6, 8)The **noble gases** are the most inert group of all. Under normal circumstances, they will not form compounds. However, in 1962, scientists made xenon (Xe, 54) compounds with ultra-reactive fluorine (F, 9).

Periodic Trends

- Atomic Radius: distance from nucleus to outermost electron; increases as one moves down a period, and from right to left. These are due to more electron shells and lower electronegativity - except in the case of the noble gasses, which have enough electrons already and so have no electronegativity, but the protons have enough electronegativity to pull electrons in.
- Ionization Energy: amount of energy required to move one electron from an atom; increases as one moves up a period, and from left to right. Neutral to negative/ gives away electron/ endothermic reaction.
- Redox properties: the possibility the element will be involved in a redox equation; increases as one moves outward from the center of the table, with the exception of the inert noble gasses.

Bonding Trends The location of an element on the Periodic Table governs the element's bonding behaviors. For example, sodium (Na, 23) is an alkali metal, so it prefers to give away its outermost electron and form salts, while neon (Ne, 10), being a noble gas, refuses to bond due to its full outer shell.

Ionic vs. Covalent An **ionic bond** is a chemical bond where one ion loses one or more electrons and transfers them to another ion. These are generally seen in metal-nonmetal combinations, such as sodium chloride (NaCl), in which an electron is transferred from the sodium atom to the chlorine atom. A **covalent bond** is a chemical bond where one or more electrons from one element leave that element and fill the outermost electron shell of the other element. In most simple cases, this results in both elements having full outer shells of electrons. These are usually seen in nonmetal-nonmetal bonds, such as in water (H2O), in which the electrons are shared between the hydrogen atoms and the oxygen atom.

Ionic Charges If an atom gains or loses electrons, it is known as an **ion.** When it loses electrons, it has a positive charge and is known as a **cation.** When an ion gains electrons, it has a negative charge and is called an **anion.**

Polar covalent bonds When the atoms in a covalent bond do not share the electrons equally, it is called a **polar covalent bond.** For example, in water (H2O), the oxygen atom tends to tug on the shared electrons more than the hydrogen atoms, thus creating a slightly negative charge around the oxygen atom.

redox reaction, or an oxidation/reduction reaction, occurs when one reactant is **oxidized**, or loses electrons, and one reactant is **reduced**, or gains electrons. A simple way to tell the difference is OIL RIG (Oxidation Is Losing; Reducing Is Gaining) or LEO says GER (Lose Electrons - Oxidize; Gain Electrons - Reduce). The **oxidizing agent** is reduced, and the **reducing agent** is oxidized. A **half-reaction** is exactly what it sounds like - half a reaction. It focuses exclusively on one portion of the reaction, either oxidation or reduction.

oxidation number of an atom in a molecule refers to the electrons being shared. If, in a covalent bond, it is basically "giving" an electron to the other atom, then that counts as a +1 on it's oxidation number. If, in a covalent bond, it is "taking" an electron from the other atom, then that counts as a -1 on its oxidation number. Something that is important to realize is that it is not actually giving or taking an electron, that only occurs in ionic bonds, but that it or the other atom is getting the electron more often. When something is oxidized its oxidation number increases, and when something is reduced it's oxidation number decreases. The **octet rule** states that all atoms (excluding H and He) are trying to get an octet (8) electrons. H is trying to get either 2 or 0 electrons, and He already has 2.

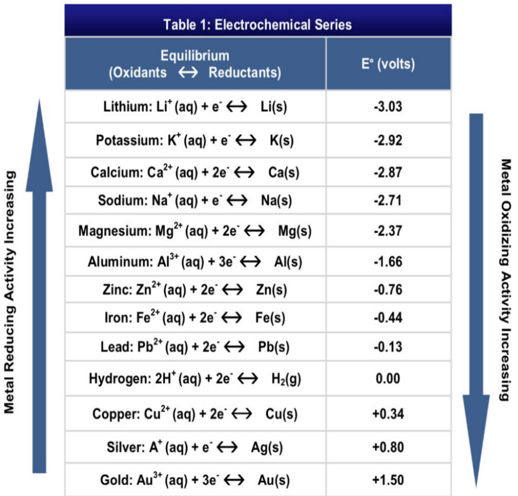


Table 1: Electrochemical Series

Equilibrium (Oxidants ↔ Reductants)	E° (volts)
Lithium: $Li^+(aq) + e^- \leftrightarrow Li(s)$	-3.03
Potassium: $K^+(aq) + e^- \leftrightarrow K(s)$	-2.92
Calcium: $Ca^{2+}(aq) + 2e^- \leftrightarrow Ca(s)$	-2.87
Sodium: $Na^+(aq) + e^- \leftrightarrow Na(s)$	-2.71
Magnesium: $Mg^{2+}(aq) + 2e^- \leftrightarrow Mg(s)$	-2.37
Aluminum: $Al^{3+}(aq) + 3e^- \leftrightarrow Al(s)$	-1.66
Zinc: $Zn^{2+}(aq) + 2e^- \leftrightarrow Zn(s)$	-0.76
Iron: $Fe^{2+}(aq) + 2e^- \leftrightarrow Fe(s)$	-0.44
Lead: $Pb^{2+}(aq) + 2e^- \leftrightarrow Pb(s)$	-0.13
Hydrogen: $2H^+(aq) + 2e^- \leftrightarrow H_2(g)$	0.00
Copper: $Cu^{2+}(aq) + 2e^- \leftrightarrow Cu(s)$	+0.34
Silver: $A^+(aq) + e^- \leftrightarrow Ag(s)$	+0.80
Gold: $Au^{3+}(aq) + 3e^- \leftrightarrow Au(s)$	+1.50

Figuring Out Oxidation Numbers

1. The most electronegative atom gets the electron.

2. The oxidation numbers add up to the charge of the molecule. If you have a charge put it on the atoms which have not reached an octet of electrons. Change their oxidation number accordingly.

Take the example of ethanol, CH3CH2OH Number the C atoms. There is one on the left that will be called C1. There is another one in the middle that will be called C2. C1 has 3 bonds with H atoms. In each bond C1 is the more electronegative. So each of the 3 H atoms have an oxidation number of +1. So far C1's oxidation number is -3. C1 has another bond that is with C2. They each have the same electronegativity, so there is no change in their oxidation numbers. Thus, C1 has an oxidation number of -3. C2 so far has an oxidation number of 0 from the C1-C2 bond. It is also bonded to 2 H atoms, giving it an oxidation number of -2 and giving each of the H atoms an oxidation number of +1. Lastly, it has a bond with the O atom. O is more electronegative, so C2's oxidation number is -1. O so far has an oxidation number of -1 from the C-O bond. It is also bonded with an H atom. Since it is more electronegative, it gets the electron. Thus, the H atom has an oxidation number of +1, and O has an oxidation number of -2. Last but not least, add up all the oxidation numbers. All 6 of the H atoms have an oxidation number of +1. C1 has an oxidation number of -3. C2 has an oxidation number of -1. O has an oxidation number of -2. 6(+1) -3 -1 -2 = 0 This is the charge of ethanol, so the procedure has been completed.

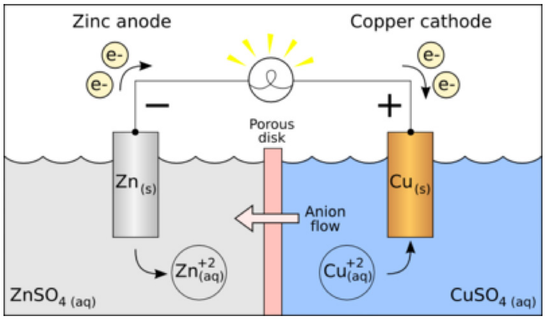
Balancing Oxidation/Reduction Reactions

1. **Split it into 2 half-reactions.** Usually, the basis of a redox reaction will be given with two compounds reacting to form two more compounds. Aside from oxygen and hydrogen, each reactant will have a corresponding element with a product. These two compounds will form one half-reaction; the other two form the other. Balance the coefficients of key compounds if necessary, i.e. Balancing Cr with Cr2O7. ***Don't worry about balancing extra hydrogens or oxygens yet.***
2. **Balance all non-hydrogen or oxygen elements**. You may have balanced the key components of the half-reaction, but sometimes you'll have a pesky oxygen or hydrogen messing things up. To get this to balance, you must add ions to the other side to balance the reaction out. This process differs for acidic vs. basic solutions. Do **NOT** try to balance one equation with the other, this step **ONLY** comes at the end.
3. If it's in an **acidic medium**:
    1. **Balance oxygen** by adding H2O to the appropriate side. Add a coefficient to this H2O if necessary.
    2. **Balance hydrogen** by adding H+ to the appropriate side. This can be on either side, depending on how many H2O's you added. Add a coefficient to this if necessary.
    3. **Balance the charge** by adding electrons. Now you've balanced the equation in terms of elements, but the charge may not be balanced yet. To balance this, add electrons to the side with a higher charge until the total charge of each half of the half-reaction is the same.

4. If it's a **basic medium**
    1. **Balance oxygen** as above.
    2. **Balance hydrogen**. Instead of balancing this with H+, you need to balance it with OH-. This means that you may get extra oxygens. If this happens, add another H2O on the other side and continue adding OH- until it balances. There is also **another method** that is detailed below.
    3. **Balance the charge** as above.
5. Now, we can add the reactions together to come up with our final reaction. **Multiply each reaction by an integer** so that there are the same number of electrons on each side (i.e. they cancel out). This means that the electrons of one half-reaction should be on the OPPOSITE side of the electrons in the other half-reaction. If this is not the case, go back and check your work. More than likely, there's a mistake in there somewhere.
6. Combine the half-reactions and cancel. The electrons should cancel out completely, and H2O's and H+'s may cancel somewhat. If you have time, it's usually a good idea to make sure that the equation is balanced by elements and by charge.

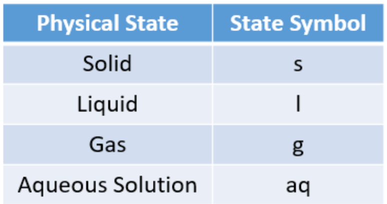
- **Second method for balancing redox reactions in basic solution**: If it's in a basic medium, add OH- to each side of the final equation until all H+ is gone; then, cancel again. Remember that OH- + H+ ---> H2O in this step.

Activity Series One task participants may be asked to complete in this event is to construct an activity series based on what ions react with others. This activity series dictates what elements oxidize more easily than others. One of the most common ways that you will have to make an activity series in Chem Lab is through performing single replacement reactions by putting metal strips in a metal solution and seeing if there is a reaction. Each team will have a set of solutions and metals, and will have to perform each possible combination of metal to solution. A table can be formed recording which combinations result in reactions. The metal that reacts the most is the one that oxidizes the most easily, while the metal that does not react at all is the one that reduces the most easily. Once complete, the activity series should look similar to the one at right. If not, you likely made a mistake and should recheck your work.

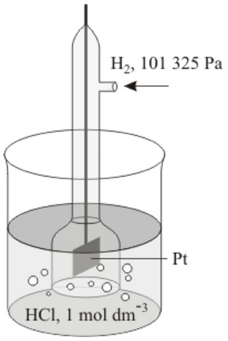


Electrochemical and Voltaic Cells Electrochemical cells results in an exchange of electrons in a redox reaction. There are two main types of electrochemical cells. A voltaic, or galvanic, cell is composed of two metals connected by a salt bridge. It uses the electron exchange to generate the current. It consists of two half-cells, each of which contains a metal solution with that metal submerged in it. The cell in which oxidation occurs is called the **anode**, and the cell in which reduction occurs is called the **cathode**. You can remember this by knowing that reduction has a "c" in it, and cathode starts with a "c". Electrons flow from the anode to the cathode. Another, more sure fire way of telling where oxidation and reduction occurs is as so: the cathode is positive (you can remember this by "cats are positive", even if they're not). Thus, electrons flow **to** the cathode, meaning that reduction is occurring at the cathode. This means that oxidation occurs at the anode. A porous barrier allows ions to flow from the anode compartment to the cathode compartment and vice versa, balancing the charge. The electrode compartments are called **half-cells**. Salt bridges may be used as an alternative to porous barriers.

Electrolytic Cells Electrolytic cells use a current to decompose chemical compounds. One principle use of electrolytic cells is to electroplate objects such as nails and silverware. **Coulombs** is a measure of charge. **Current** is a measure of the flow of electrons. The SI unit of current is the ampere, expressed as C/s or coulombs per second. Coulombs = amperes × seconds Given the current run through an electrolytic cell and the time it is run, you can calculate the number of coulombs. There are 1.602E-19 coulombs in an electron. From the amount of coulombs you may calculate the number of electrons used to reduce. Say that you are electroplating copper onto a plate. Cu2++2e−→Cu Given the number of electrons used to reduce the copper ions, you may calculate the amount of Cu electroplated onto the plate. This is applicable to any electroplating situation.



Physical State	State Symbol
Solid	s
Liquid	l
Gas	g
Aqueous Solution	aq



Electron Potential There are two different analogies for understanding electron potential or **voltage**. One is water. Electron potential corresponds to the water pressure. The higher the pressure, the stronger the stream that flows. Electron potential does not correspond to the strength of the stream, since different sized pipes with the same water pressure will have different strength streams. The second analogy is height. Higher electron potential corresponds to higher height. From higher height you can drop, while doing work, to lower height. Something to note is that electron potential is not absolute, it is with respect to. Standing on a 10 ft. high cliff and dropping a ball is the same as standing on the edge of a 10 ft. deep pit and dropping a ball (ignoring changes in gravity: analogies are not perfect). It is the same way with electron potential. You must define a zero before you can say what the electron potential is. Because of this it is quite possible to have negative electron potential.

Electromotive Force (emf) The **emf** of a cell, measured in volts, is the potential difference between the cathode and the anode of a cell. It tells you how much potential there is to do work. Electromotive means "causing electron motion". It is quite easy to measure emf. Take a voltmeter and touch the probes to the cathode and anode. The voltmeter will tell you what the voltage difference, or emf, is.

Standard Reduction (Half-Cell) Potentials **Reduction Potentials** tell you how much something "wants" to reduce. For example, Cu2+ (with a reduction reaction of Cu2++2e−→Cu) has a higher reduction potential then Fe2+ (with a reduction reaction of Fe2++2e−→Fe). This means that Cu2+ "wants" to reduce more then Fe2+ Like all potentials, reduction potentials are not absolute and have to be with respect to something.

Reduction Half-Reaction	Potential(V)
F2+2e−→2F−	+2.87
MnO4−+8H++5e−→Mn2++4H2O	+1.51
Cl2+2e−→2Cl−	+1.36
Cr2O72−+14H++6e−→2Cr3++7H2O	+1.33
O2+4H++4e−→2H2O	+1.23
Br2+2e−→2Br−	+1.06
NO3−+4H++3e−→NO+2H2O	+0.96
Ag++e−→Ag	+0.80
Fe3++e−→Fe2+	+0.77
O2+2H++2e−→H2O2	+0.68
MnO4−+2H2O+3e−→MnO2+4OH−	+0.59

in

I2+2e−→2I−	+0.54
O2+2H2O+4e−→4OH−	+0.40
Cu2++2e−→Cu	+0.34
2H++2e−→H2	0[defined]
Ni2++2e−→Ni	-0.28
Fe2++2e−→Fe	-0.44
Zn2++2e−→Zn	-0.76
2H2O+2e−→H2+2OH−	-0.83
Al3++3e−→Al	-1.66
Na++e−→Na	-2.71
Li++e−→Li	-3.05

standard reduction potentials:

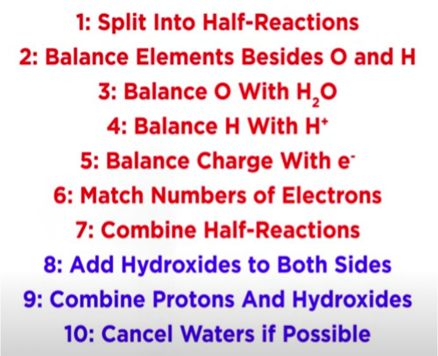
Combustion reactions are reduction-oxidation reactions (also called *redox* reactions) involving oxygen or another oxidizing agent. Combustion reactions are also exothermic reactions since they give off heat. The most common combustion reactions form carbon dioxide, water, and energy. For example, here is the combustion reaction for methane:

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ + energy

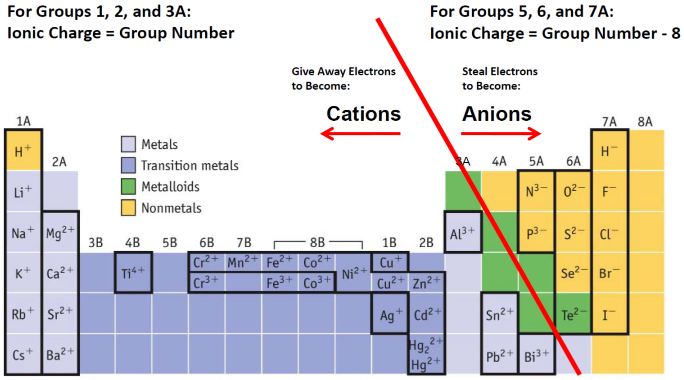
Combustion can also occur with nitrogen instead of carbon. There is also combustion with only hydrogen and oxygen, and this case only water forms as product.

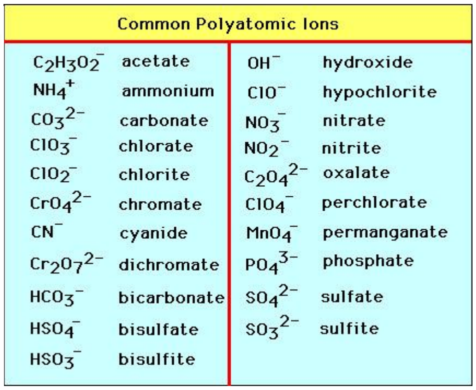
$2H_2 + O_2 \rightarrow 2H_2O$ + energy

Since combustion reactions are among the most common exothermic reactions, it is a good idea to know the combustion reactions of several important compounds, or at least know how to go about finding it quickly. Since all of them have a similar form, you can guess what the products will be, which make it easier.



1: Split Into Half-Reactions
2: Balance Elements Besides O and H
3: Balance O With H_2O
4: Balance H With H^+
5: Balance Charge With e^-
6: Match Numbers of Electrons
7: Combine Half-Reactions
8: Add Hydroxides to Both Sides
9: Combine Protons And Hydroxides
10: Cancel Waters if Possible





Common Polyatomic Ions

$C_2H_3O_2^-$	acetate	OH^-	hydroxide
NH_4^+	ammonium	ClO^-	hypochlorite
CO_3^{2-}	carbonate	NO_3^-	nitrate
ClO_3^-	chlorate	NO_2^-	nitrite
ClO_2^-	chlorite	$C_2O_4^{2-}$	oxalate
CrO_4^{2-}	chromate	ClO_4^-	perchlorate
CN^-	cyanide	MnO_4^-	permanganate
$Cr_2O_7^{2-}$	dichromate	PO_4^{3-}	phosphate
HCO_3^-	bicarbonate	SO_4^{2-}	sulfate
HSO_4^-	bisulfate	SO_3^{2-}	sulfite
HSO_3^-	bisulfite		

Some things to consider when graphing-

1. Unless you have a good reason to do otherwise, the independent data goes on the x-axis while the dependent data goes on the x-axis.
2. A graph must have a title. Something like “Y as a Function of X” or “Y Versus X” is usually a good choice. Just don’t literally use “Y” and “X” in the title!
3. Label the axes and include the units.
4. The scale for each axis must be linear although the x-axis can have a different scale than the y-axis.
5. Unless you have a good reason to not include it, show the origin (0,0) on the graph.
6. The value of the slope needs to have appropriate units and they come from the units for the y-axis divided by the units for the x-axis.
7. Unless you know otherwise, determine the number of digits for the slope by using the division rule for significant figures. My mass data has a minimum of 5 significant figures and the volume data has a minimum of four significant figures. Therefore, I’ve calculated the slope to four significant figures.