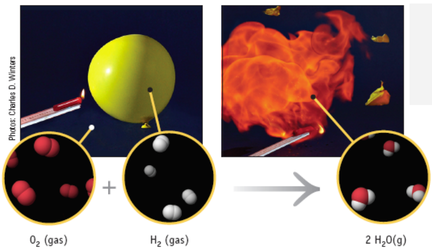
Chapter 7 Part I: Stoichiometry of Chemical Reactions *(7.1 - 7.2 only)*



Chemistry 221
Professor Michael Russell
http://mhchem.org/221

Last update: 7/7/25

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7.1 - Writing and Balancing Chemical Equations

A **balanced chemical equation** *(or reaction)* uses symbolism to represent the identities *and* the relative quantities of substances undergoing a chemical (or physical) change.

Chemical equations depict the kind of **reactants** *(left)* and **products** *(right)* and their relative amounts in a reaction.

$$4\ Al(s)\ +\ 3\ O_2(g)\ \rightarrow\ 2\ Al_2O_3(s)$$

The numbers in the front are called **stoichiometric coefficients**

The letters (**s**), (**g**), (**aq**) and (**l**) are the physical states of compounds (**solid**, **gas**, **liquid** and '**dissolved in water**' (aq))

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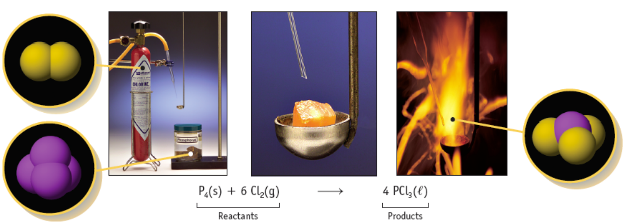
Reaction of Phosphorus with Cl_2



Notice the stoichiometric coefficients and the physical states of the reactants and products.

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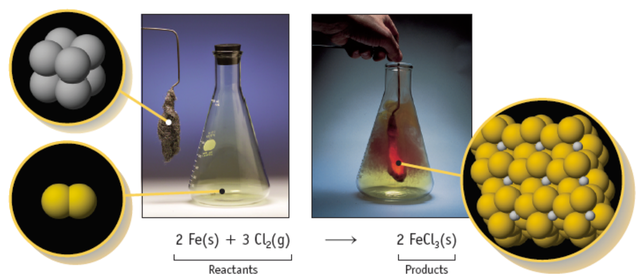
Reaction of Iron with Cl_2



Evidence of a chemical reaction:

heat change, precipitate formation, gas evolution, color change

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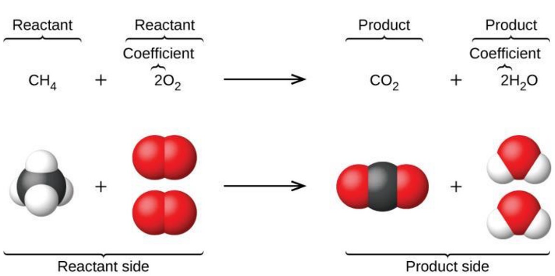
Reaction of Methane (CH_4) with O_2



Use *smallest* coefficients possible, i.e. 1:2:1:2 *(above)*, not 2:4:2:4, *etc.*

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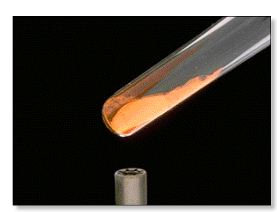
Chemical Equations

Because the same atoms are present in a reaction at the beginning and at the end, the amount of matter in a system does not change.

This is the Law of the Conservation of Matter

Also known as the Law of Mass Action



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Lavoisier

Because of the principle of the conservation of matter,

an equation must be balanced.

It must have the same number of atoms of the same kind on both sides.

Lavoisier, 1788

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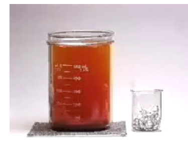
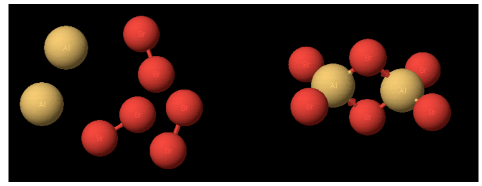
Balancing Equations



2 Al(s) + 3 Br_2(liq) ---> Al_2Br_6(s)



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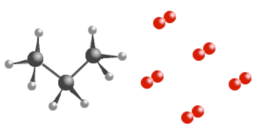
Balancing Equations



___C_3H_8(g) + _____ O_2(g) → _____CO_2(g) + _____ H_2O(g)

___B_4H_{10}(g) + _____ O_2(g) → ___ B_2O_3(g) + _____ H_2O(g)

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Balancing Equations - *hints*

A reaction which is **heated** may include an uppercase Greek letter delta (Δ) over the arrow:

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$$

It is sometimes convenient to use fractions instead of integers as intermediate coefficients.

$$C_2H_6 + {}^{7}/_{2}\, O_2 \rightarrow 3\, H_2O + 2\, CO_2 \quad \text{(balanced)}$$

Generally fractions are not used, so when balance is achieved, all the equation's coefficients are multiplied by a whole number to convert to whole numbers.

$$2\, C_2H_6 + 7\, O_2 \rightarrow 6\, H_2O + 4\, CO_2 \quad \text{(balanced)}$$

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Balancing Equations - More Hints

Balance those atoms which occur in only one compound on each side last *(i.e. O_2 in previous examples)*

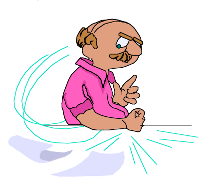
Balance the remaining atoms first. Try not to use fractions as coefficients.

Never change the subscripts on a formula, only change the coefficients

Reduce coefficients to smallest whole integers

Check your answer *if uncertain*

Check that charges are balanced



Practice, practice, practice!

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Aqueous Compounds

Consider: **$CaCl_2$(aq) + 2 $AgNO_3$(aq) → $Ca(NO_3)_2$(aq) + 2 AgCl(*s*)**

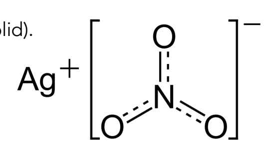
When ionic compounds dissolve in water, they dissociate into their constituent ions. *Examples:*

$CaCl_2$(aq) → Ca^{2+}(aq) + 2 Cl^{1-}(aq)

2 $AgNO_3$(aq) → 2 Ag^{1+}(aq) + 2 NO_3^{1-}(aq)

$Ca(NO_3)_2$(aq) → Ca^{2+}(aq) + 2 NO_3^{1-}(aq)

...but AgCl does not dissolve in water (it stays a solid).



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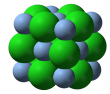
Aqueous Compounds



Consider: **$CaCl_2$(*aq*) + 2 $AgNO_3$ (*aq*) ⟶ $Ca(NO_3)_2$(*aq*) + 2 AgCl(*s*)**

We can re-write the equation as:

Ca^{2+}(aq) + 2 Cl^{1-}(aq) + 2 Ag^{1+}(aq) + 2 NO_3^{1-}(aq) ⟶ Ca^{2+}(aq) + 2 NO_3^{1-}(aq) + 2 AgCl(s)

Notice that Ca^{2+}(aq) and 2 NO_3^{1-}(aq) appear on both sides - they are **spectator ions** and can be removed from the reaction to get the **net ionic reaction**:

2 Cl^{1-}(aq) + 2 Ag^{1+}(aq) ⟶ 2 AgCl(s) *or better yet:*

Cl^{1-}(aq) + Ag^{1+}(aq) ⟶ AgCl(s)

*This is the **net ionic reaction***
***No spectator ions** in net ionic reactions*

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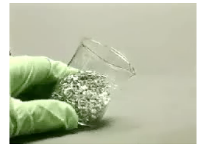
Net Ionic Equations



The two Cl^- ions are SPECTATOR IONS - they do not participate. Could have used NO_3^- (via HNO_3)

Mg(s) + 2 HCl(aq) → H_2(g) + $MgCl_2$(aq)

Break up (aq), leave gases, solids and liquids alone, to write:

Mg(s) + 2 H^+(aq) + 2 Cl^-(aq) → H_2(g) + Mg^{2+}(aq) + 2 Cl^-(aq)

Cl^- is a spectator - leave out - write the Net Ionic Equation:

Mg(s) + 2 H^+(aq) → H_2(g) + Mg^{2+}(aq)

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Net Ionic Equations

K_2CrO_4(aq) + $Pb(NO_3)_2$(aq) → $PbCrO_4$(s) + 2 KNO_3(aq)

Net Ionic Equation:

Pb^{2+}(aq) + CrO_4^{2-}(aq) → $PbCrO_4$(s)

K^+ and NO_3^- are spectators

See Net Ionic Reactions Handout

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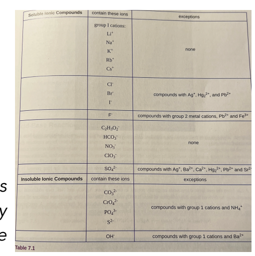
7.2 - Classifying Chemical Reactions - *Precipitation*

A **precipitation reaction** is a reaction where dissolved substances create solid products (a **precipitate**.) Chemists use a **solubility table** to know common combinations of chemicals which result in **soluble** (no solids) or **insoluble** (solids form) products.

Our solubility table is slightly different than the textbook's version, but similar.



Soluble Ionic Compounds	contain these ions	exceptions
	group I cations: Li^+ Na^+ K^+ Rb^+ Cs^+	none
	Cl^- Br^- I^-	compounds with Ag^+, Hg_2^{2+}, and Pb^{2+}
	F^-	compounds with group 2 metal cations, Pb^{2+} and Fe^{3+}
	$C_2H_3O_2^-$ HCO_3^- NO_3^- ClO_3^-	none
	SO_4^{2-}	compounds with Ag^+, Ba^{2+}, Ca^{2+}, Hg_2^{2+}, Pb^{2+} and Sr^{2+}
Insoluble Ionic Compounds	contain these ions	exceptions
	CO_3^{2-} CrO_4^{2-} PO_4^{3-} S^{2-}	compounds with group 1 cations and NH_4^+
	OH^-	compounds with group 1 cations and Ba^{2+}

Table 7.1

Textbook's Solubility Guide

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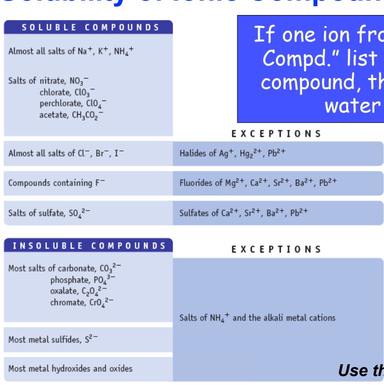
Water Solubility of Ionic Compounds



SOLUBLE COMPOUNDS	EXCEPTIONS
Almost all salts of Na^+, K^+, NH_4^+	
Salts of nitrate, NO_3^-; chlorate, ClO_3^-; perchlorate, ClO_4^-; acetate, $CH_3CO_2^-$	
Almost all salts of Cl^-, Br^-, I^-	Halides of Ag^+, Hg_2^{2+}, Pb^{2+}
Compounds containing F^-	Fluorides of Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}
Salts of sulfate, SO_4^{2-}	Sulfates of Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}

INSOLUBLE COMPOUNDS	EXCEPTIONS
Most salts of carbonate, CO_3^{2-}; phosphate, PO_4^{3-}; oxalate, $C_2O_4^{2-}$; chromate, CrO_4^{2-}	Salts of NH_4^+ and the alkali metal cations
Most metal sulfides, S^{2-}	
Most metal hydroxides and oxides	

If one ion from the "Soluble Compd." list is present in a compound, the compound is water soluble.

$Ba(NO_3)_2$(aq)? soluble

$BaCl_2$(aq)? soluble

$BaSO_4$(aq)? insoluble *should write $BaSO_4$(s)!*

Use this solubility guide in CH 221-223!

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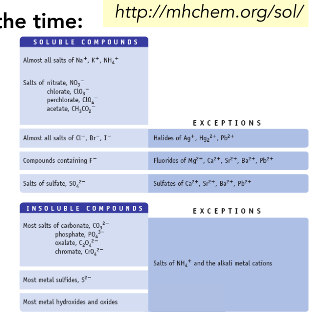
Water Solubility Of Ionic Compounds

Many ionic compounds dissolve in water (soluble), but many are insoluble and do not dissolve.

Many ions make compounds soluble *all* of the time:

Solubility guide: http://mhchem.org/sol/

***Examples:* Na^+, K^+, Li^+, NH_4^+, NO_3^-, ClO_3^-, ClO_4^-, $CH_3CO_2^-$, and *most* SO_4^{2-}, Cl^-, Br^- and I^- compounds.**



SOLUBLE COMPOUNDS	EXCEPTIONS
Almost all salts of Na^+, K^+, NH_4^+	
Salts of nitrate, NO_3^-; chlorate, ClO_3^-; perchlorate, ClO_4^-; acetate, $CH_3CO_2^-$	
Almost all salts of Cl^-, Br^-, I^-	Halides of Ag^+, Hg_2^{2+}, Pb^{2+}
Compounds containing F^-	Fluorides of Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}
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INSOLUBLE COMPOUNDS	EXCEPTIONS
Most salts of carbonate, CO_3^{2-}; phosphate, PO_4^{3-}; oxalate, $C_2O_4^{2-}$; chromate, CrO_4^{2-}	Salts of NH_4^+ and the alkali metal cations
Most metal sulfides, S^{2-}	
Most metal hydroxides and oxides	

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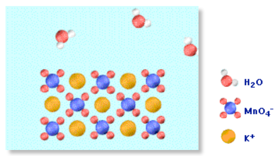
Ionic Compounds in Aqueous Solution

Soluble ionic compounds dissolve in water to make aqueous solutions.

Insoluble ionic compounds stay undissolved in water.

$KMnO_4$ in water



$K^+(aq)$ + $MnO_4^-(aq)$

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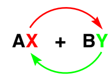
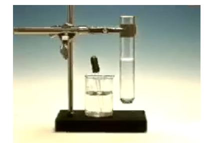
Chemical Reactions In Water

Many reaction types are Exchange Reactions



AX + BY → AY + BX

The anions exchange places between cations



$Pb(NO_3)_2(aq) + 2\ KI(aq)$
$\rightarrow PbI_2(s) + 2\ KNO_3(aq)$

Exchange reactions often called Double Displacement Reactions

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Precipitation Reactions

A precipitation reaction is an *exchange reaction* where a precipitate (i.e. solid) is formed as a product

The "driving force" is the formation of an insoluble compound - a precipitate.

$Fe(NO_3)_3(aq) + 3\ NaOH(aq) \rightarrow$
$3\ NaNO_3(aq) + Fe(OH)_3(s)$

Net ionic equation

$Fe^{3+}(aq) + 3\ OH^-(aq) \rightarrow Fe(OH)_3(s)$

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Acid-Base Reactions

An acid-base reaction is an *exchange reaction* where an acid and a base create water and a 'salt'

Acids react readily with bases. The "driving force" is the formation of water.

$NaOH(aq) + HCl(aq) \rightarrow$
$NaCl(aq) + H_2O(liq)$

Net ionic equation:

$OH^-(aq) + H^+(aq) \rightarrow H_2O(liq)$

This applies to ALL reactions of STRONG acids and bases.

Acid-base reactions often called "neutralizations", water and "salt" created

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What is an Acid?

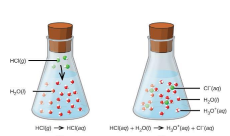
Hydronium often written as just H^+

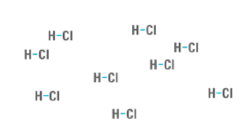
An **acid** is a substance which dissolves in water to create the **hydronium ion**, H_3O^+.

Acids that give the maximum amount of hydronium possible are called **strong acids**.

There are five strong acids *(important):*

HCl	hydrochloric acid
HBr	hydrobromic acid
HI	hydroiodic acid
HNO_3	nitric acid
$HClO_4$	perchloric acid





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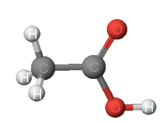
Weak Acids



acetic acid

Most acids do not create the maximum amount of hydronium possible; they are called **weak acids**.

Most foods are weak acids!

Assume an acid is weak unless you know it is strong

Examples:

CH_3CO_2H	acetic acid
H_2CO_3	carbonic acid
H_3PO_4	phosphoric acid
HF	hydrofluoric acid

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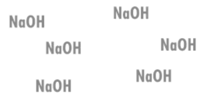
What is a Base?

A **base** is a metal hydroxide. Hydroxide ions react with hydronium ions to create water (and energy!)
Bases that give the maximum amount of hydroxide possible are called **strong bases**.

There are three strong bases *(important):*

NaOH	**sodium hydroxide**
KOH	**potassium hydroxide**
LiOH	**lithium hydroxide**



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Weak Bases

Most bases do not create the maximum amount of hydroxide possible; they are called **weak bases**.
Most household cleaners are weak bases
Assume a base is weak unless you know it is strong

Examples:

NH_3	**ammonia**
Na_2CO_3	**sodium carbonate**
Na_3PO_4	**sodium phosphate**
N_2H_4	**hydrazine**

ammonia creates ammonium hydroxide in water

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Common Acids and Bases

Strong Acids (Strong Electrolytes)		Strong Bases (Strong Electrolytes)	
HCl	Hydrochloric acid	LiOH	Lithium hydroxide
HBr	Hydrobromic acid	NaOH	Sodium hydroxide
HI	Hydroiodic acid	KOH	Potassium hydroxide
HNO_3	Nitric acid		
$HClO_4$	Perchloric acid		
Weak Acids (Weak Electrolytes)*		**Weak Base (Weak Electrolyte)**	
H_3PO_4	Phosphoric acid	NH_3	Ammonia
H_2CO_3	Carbonic acid		
CH_3CO_2H	Acetic acid		
$H_2C_2O_4$	Oxalic acid		
$C_4H_6O_6$	Tartaric acid		
$C_6H_8O_7$	Citric acid		
$C_9H_8O_4$	Aspirin		

*These are representative of hundreds of weak acids.

Know the strong acids & bases!

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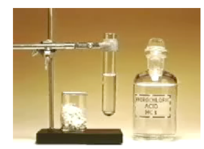
Gas-Forming Reactions



A gas forming reaction is an *exchange reaction* where one of the products is a gas. Gases often result from unstable products that break down into a gas immediately after the chemicals are mixed. *Example:*

$$CaCO_3(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + H_2CO_3(aq)$$

Carbonic acid *(H_2CO_3)* **is unstable** and forms **CO_2 & H_2O**

$$H_2CO_3(aq) \rightarrow CO_2(g) + H_2O(l)$$

so we should write:

$$CaCO_3(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

Another gas forming species: **ammonium hydroxide**

$$NH_4OH(aq) \rightarrow NH_3(g) + H_2O(l)$$

Be on the lookout for H_2CO_3 and NH_4OH as products! They will break down!

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Combustion Reactions

A special example of a gas-forming reaction; not an exchange reaction. Used in quantitative chemistry; high temperatures

***Reactants:** (often) oxygen (O_2) and "something organic" (C, H, sometimes O or N)*

***Products:** water and carbon dioxide (also NO_2 if N present)*

Examples:

$$C_2H_4(g) + 3\ O_2(g) \rightarrow 2\ H_2O(g) + 2\ CO_2(g)$$

$$4\ C_6H_5NO_2(g) + 29\ O_2(g) \rightarrow 10\ H_2O(g) + 24\ CO_2(g) + 4\ NO_2(g)$$

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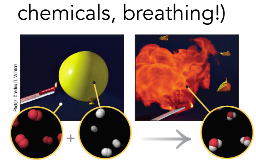
Oxidation-Reduction Reactions

An **oxidation-reduction reaction** (often called a **redox** reaction) is a reaction where electrons are transferred from one reactant (the **reducing agent**) to another reactant (the **oxidizing agent**).
Knowing where electrons come from (and go to) important!
Very powerful and common reaction type (batteries, making chemicals, breathing!)



$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(liq)$$

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LEO says GER

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LEO says GER

Lose **E**lectrons **O**xidized — **G**ain **E**lectrons **R**educed

$Zn(s) \rightarrow Zn^{2+} + 2e^-$ — ***Oxidized*** (Zn is the *reducing agent*)

$Cu^{2+} + 2e^- \rightarrow Cu(s)$ — ***Reduced*** (Cu^{2+} is the *oxidizing agent*)

*Can also use **"OIL RIG":***
***OIL = "Oxidation is Losing"** (electrons)*
***RIG = "Reduction is Gaining"** (electrons)*

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Redox Reactions

Lose **E**lectrons **O**xidized — **G**ain **E**lectrons **R**educed

In all reactions: if something has been oxidized then something has also been reduced:

$Cu(s) + 2\ Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2\ Ag(s)$

Oxidized species become more positive (lose electrons), reduced species become more negative (gain electrons)

Oxidation numbers help visualize electron transfer pathways

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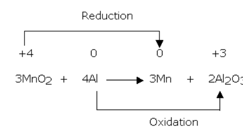
Oxidation Numbers

*Use **oxidation number rules** to determine redox activity:*

- **Atoms in free element have ox. no. = 0**
  $Zn(s)$, $O_2(g)$, $Br_2(liq)$
- *In simple ions,* **ox. no. = charge on ion**
  -1 for Cl^-, +2 for Mg^{2+}
- **In compounds, F is *always* -1, O is -2** *(except peroxides (O = -1) and with F) and* **H is +1** *(except hydrides (H = -1))*
- **Sum of oxidation numbers = 0** *for a* **compound or equals the overall charge** *for an* **ion**



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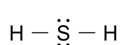
Oxidation Numbers

Example: Find the oxidation number for each element in **H_2S**.

The oxidation number for **H is +1**, H_2S is neutral, so:

Charge on H_2S = 0 = (2 × 1) + 1x

x = −2 *the oxidation number for S is -2!*

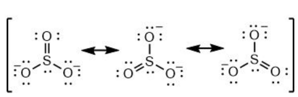


Example: Find the oxidation number for each element in **sulfite, SO_3^{2-}**

The oxidation number for **O is -2**, sulfite has a -2 charge, so:

-2 = 1x + (3 × -2)

x = +4 *the oxidation number for S is +4!*



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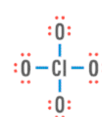
Oxidation Numbers

Determining oxidation numbers takes practice

HF
H: +1
F: -1



ClO_4^-
Cl: +7
O: -2

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Single Replacement Reactions

Single Replacement reactions are *always* redox reactions

Examples:

$Cu(s) + 2 AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

$Mg(s) + 2 HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

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Redox Reaction Examples

NO = reducing agent
O_2 = oxidizing agent
$2 NO + O_2 \rightarrow 2 NO_2$

Fe = reducing agent
Cl_2 = oxidizing agent
$2 Fe + 3 Cl_2 \rightarrow 2 FeCl_3$

reducing agent = oxidized
oxidizing agent = reduced

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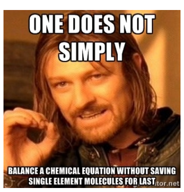
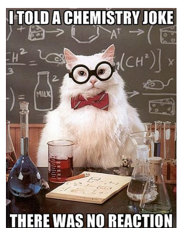
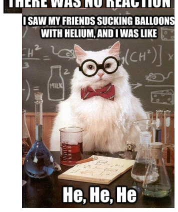
End of Chapter 7 Part 1

See also:

- *Chapter Seven Part 1 Study Guide*
- *Chapter Seven Part 1 Concept Guide*
- *Important Equations (following this slide)*
- *End of Chapter Problems (following this slide)*







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Important Equations, Constants, and Handouts from this Chapter:

- Know how the **solubility guide** works
- Know what makes an acid acidic (and bases basic) and strong or weak; know how to use the pH scale
- Know how to write and determine net ionic equations and find spectator ions
- Know how to use molarity with solution stoichiometry problems

Balancing Equations: Reactants, Products, states of matter (s, l, g, aq), stoichiometric coefficients, Law of Conservation of Matter ("mass action")

Solutions: Solute, solvent, aqueous, solubility (use the Net Ionics solubility table), precipitation, types of reactions, molarity (M)

Know the five types of reactions: precipitation, acid-base, gas forming, combustion and redox. Know how to determine if something has been **oxidized** or **reduced** (and the **oxidizing agent** and **reducing agent**)

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End of Chapter Problems: *Test Yourself*

1. Predict whether these compounds would be labeled as insoluble or soluble: HCl, NaCl, AgCl
2. Predict the products of this precipitation reaction and write the net ionic equation: $NiCl_2(aq) + (NH_4)_2S(aq) \rightarrow$? List any spectator ions.
3. In the following reaction, decide which reactant is oxidized and which is reduced. Designate the oxidizing agent and the reducing agent. **$Si(s) + 2 Cl_2(g) \rightarrow SiCl_4(l)$**
4. Identify the ions and their concentration that exist in this aqueous solution: **0.25 M $(NH_4)_2SO_4$**

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End of Chapter Problems: *Answers*

1. Soluble: HCl(aq), NaCl(aq). Insoluble: AgCl(s)
2. $NiCl_2(aq) + (NH_4)_2S(aq) \rightarrow NiS(s) + 2 NH_4Cl(aq)$
   $Ni^{2+}(aq) + S^{2-}(aq) \rightarrow NiS(s)$ *Spectator ions:* NH_4^{+1} and Cl^{-1}
3. Si is oxidized and is the reducing agent; Cl_2 is reduced and is the oxidizing agent
4. 0.50 M NH_4^{+1}; 0.25 M SO_4^{2-}

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