Nuggets: Atomic Structure; Isotopes; Counting neutrons, protons, electrons; Average Atomic Mass; Ionic versus molecular compounds; Naming chemicals; Writing formulas from names; Polyatomic ions; Coulomb’s law

EXPERIMENTS

Rutherford’s Au Foil Experiment: Mass of atom is concentrated in a highly dense nucleus that was positively charged; the rest of the atom has only a little mass

ATOMIC STRUCTURE

Nucleus: contains nearly all the atom’s mass; neutrons and protons located in the nucleus; very small volume of atom; extremely dense; atom is mostly empty space

Elements: determined by the number of protons they contain

Masses of all atoms are compared to $^{12}\text{C}$ which is the standard; $^{12}\text{C}$ is defined as 12.0000... unified atomic mass units (exact value) = 12.0000...u; unified atomic mass units = u;

- 1 u = $\frac{1}{6.022 \times 10^{23}} = 1.6605 \times 10^{-24}$ g; (note: $u$ is sometimes written as amu = atomic mass units)
- Neutron (n): has a neutral charge; mass = 1u ≈ 1 $\times 10^{-24}$ g (1.6749 $\times 10^{-24}$ g)
- Proton (p$^+$): has a charge of +1; mass = 1u ≈ 1 $\times 10^{-24}$ g (1.6726 $\times 10^{-24}$ g)
- Electron (e$^-$): has a charge of -1; mass = 0.0005u ≈ 1 $\times 10^{-27}$ g (9.1094 $\times 10^{-28}$ g)

Relative masses: n > p$^+$ ≫ e$^-$; mass of n ≈ p$^+$ ≈ 1 with e$^-$ ≈ 1/2000

ISOTOPES: elements with the same #p$^+$ (same element) but different #n (different mass)

Isotope Symbol: $^A_Z\text{Element Symbol}$  $A = \text{mass number} = #\text{p}^+ + #\text{n}$; $Z = \text{atomic number} = #\text{p}^+$

C isotopes: $^{12}_6\text{C}$, $^{13}_6\text{C}$, and $^{14}_6\text{C}$; the 3 isotopes of C each have $Z = 6$ ($6p^+$) but different #n (6, 7, and 8, respectively)

AVERAGE ATOMIC MASS (AAM): weighted average of all isotopes of one element taking into account the abundance of each isotope; units = u; e.g., average atomic mass Br = 79.90u

Abundance: percent of one isotope as compared to all atoms of that given element

AVERAGE ATOMIC MASS (AAM) = $\Sigma$ (mass of isotope) x (fractional abundance)

AAM = (mass$\text{iso}_1$)(fractional abundance$\text{iso}_1$) + (mass$\text{iso}_2$)(fractional abundance$\text{iso}_2$) + ... =

$\text{AAM} = (M_1)(FA_1) + (M_2)(FA_2) + ...$

Sum of fractional abundances = 1 (i.e., 100%); for elements with only 2 isotopes: $\text{FA}_1 + \text{FA}_2 = 1$

Example 1: Calculate the average atomic mass of an element given following data: isotope 1: 27.977u, 92.21% abundance; isotope 2: 28.976u, 4.70% abundance; and isotope 3: 29.974u, 3.09% abundance.

Answer 1: AAM = (M$^1$)(FA$^1$) + (M$^2$)(FA$^2$) + (M$^3$)(FA$^3$) = (27.977)(0.9221) + (28.976)(0.0470) + (29.974)(0.0309) = 28.09u

Example 2: What are the abundances of the 2 B isotopes ($^{10}\text{B}$ mass = 10.0129u, $^{11}\text{B}$ mass = 11.0093u), if the average atomic mass of B is 10.811u?

Answer 2: AAM = (M$^1$)(FA$^1$) + (M$^2$)(FA$^2$); 10.811 = 10.0129x + 11.0093y; since there are 2 unknowns 2 equations are needed;

second equation: FA$^1$ + FA$^2$ = 1 (which is 100%); x + y = 1 → y = 1 - x; substitute this into the first equation:

10.811 = 10.0129x + 11.0093(1-x); solve for x: 10.811 = 10.0129x +11.0093x = 0.983 = 0.9964x → x = 0.1990 = 19.90%;

y = 1-x = 1-0.1990 = 0.8010 = 80.10%; since x was the abundance for $^{10}\text{B}$ → abundance of $^{10}\text{B}$ = 19.90% and abundance of $^{11}\text{B}$ = y = 80.10%}
**IONIC COMPOUNDS:** contain metal + nonmetal (can substitute a polyatomic ion for either or both)

Properties: extended solids; not discrete molecules; high melting/boiling points; conducts electricity in molten state; Examples: NaCl(s), Ca(NO₃)₂(s), NH₄NO₃(s), AgCl(s)

**Ions:** atoms or molecules that have lost/gained e⁻ (charges are not changed by gaining/losing protons!)

- **Cations:** positively charged atoms (electrons have been lost)
- **Anions:** negatively charged atoms (electrons have been gained)

**Polyatomic Ions:** “many-atom ions”; these compounds are not broken up and are treated as a group:
- NH₄⁺, H₂O⁺, OH⁻, CH₃COO⁻, NO₃⁻, MnO₄⁻, HCO₃⁻, CO₃⁻², HSO₄⁻, SO₄⁻², PO₄³⁻, ClO₄⁻, CN⁻

**Charges on elements in Ionic Compounds** based on the column of Periodic Table the element resides in:

<table>
<thead>
<tr>
<th>Group IA: +1 (H⁺, Li⁺, ...</th>
<th>Group IIA: +2 (Be²⁺, Mg²⁺, ...</th>
<th>Group IIIA: +3 (B³⁺, Al³⁺, ...</th>
<th>Group VIIA: -1 (F⁻, Cl⁻, ...</th>
</tr>
</thead>
<tbody>
<tr>
<td>Group VA: -3 (N⁻³, P⁻³, ...</td>
<td>Group VIA: -2 (O⁻², S⁻², ...</td>
<td>Group VIIA: +1 (Na⁺, K⁺, ...</td>
<td>Group VIB: +2 (Ca²⁺, Sr²⁺, ...</td>
</tr>
</tbody>
</table>

**Naming**

- **Metal + Nonmetal → ionic compound; metal name followed by nonmetal “root”+“ide”**

  *Example 3:* What is the name for CaCl₂?  →  **Answer 5:** calcium chloride

- **Contains one or two Polyatomic Ions → ionic compound; use polyatomic ion name**

  *Example 4:* What is the name for Ca₃(PO₄)₂?  →  **Answer 6:** calcium phosphate

- **Contains Transition Metal → ionic compound; metal name (charge in Roman numerals)**

  *Example 5:* What is the name for (NH₄)₂SO₄?  →  **Answer 7:** ammonium sulfate

**Writing Formulas**

1. Take metal name or polyatomic ion and write formula including charge
   - If metal has a Roman numeral, then charge is determined from that numeral (e.g., copper(II) = Cu²⁺)
   - If it does not have a Roman numeral, it is a memorized polyatomic ion (e.g., ammonium = NH₄⁺) or determined from the column in the Periodic Table it resides in (e.g., magnesium = Mg²⁺ since it is in column IIA)

2. Take nonmetal name or polyatomic ion and write formula including charge
   - If nonmetal name ends with “ide” it is a single atom ion with a negative charge (e.g., hydroxide = OH⁻) and cyanide (CN⁻) are polyatomic ions with a name that ends in “ide”
   - If nonmetal ends with “ite” or “ate”, or starts with “hypo” or “per” it is a memorized polyatomic ion

3. Balance charges within formula

*Example 7:* Write formula for chromium(III) carbonate.

  **Answer 7:** charge on Cr is → Cr³⁺ from Roman numeral; carbonate = CO₃⁻² (memorized) → Crₓ(CO₃)ᵧ; choose x and y to balance the charge: x = 2 → [2(+3) = +6] and y = 3 → [3(-2) = -6]; +6 balances -6 → Cr₂(CO₃)₃

**MOLECULAR COMPOUNDS:** contain 2 nonmetal elements;

Properties: low melting/boiling points; discrete molecules; do not conduct electricity

Examples: CO(g), CO₂(g), N₂O₄(g)

**Naming:** Prefix (if greater than 1)+1st nonmetal name & Prefix(always)+2nd nonmetal root+ “ide”

- **Prefixes:** mono (1), di (2), tri (3), tetra (4), penta (5), hexa (6), hepta (7), octa (8), nona (9), deca (10)

- **Example 8:** What is the name for NO?  (Has 2 nonmetals → molecular compound)

  **Answer 8:** Has one N → no prefix → **nitrogen**; has one O → prefix always used = mono; **monoxide** → answer: nitrogen monoxide

- **Example 9:** What is the name for N₂O?  (Has 2 nonmetals → molecular compound)

  **Answer 9:** Has two N → use prefix = di → **dinitrogen**; has one O → prefix always used = mono; **monoxide** → dinitrogen monoxide

**Writing Formulas** - translate name with prefixes

*Example 10:* What is the formula for disulfur trioxide?

*Answer 10:* disulfur → 2S; trioxide → 3O; answer: S₂O₃
ACIDS

**Naming** – chemicals named as acids are neutral compounds (e.g., \( \text{SO}_4^{2-} \) and \( \text{HSO}_4^- \) are not named as acids but \( \text{H}_2\text{SO}_4 \) is named as an acid)

**Acids that contain anions ending in**

“ide”: change “ide” to “ic”, add “hydro” in front and “acid” at the end; e.g., \( \text{Cl}^- \) = chloride; \( \text{HCl} = \text{hydrochloric acid} \)

“ate”: change “ate” to “ic” and add “acid” at the end; e.g., \( \text{SO}_4^{2-} \) = sulfate; \( \text{H}_2\text{SO}_4 \) = sulfuric acid

“ite”: change “ite” to “ous” and add “acid” at the end; e.g., \( \text{NO}_2^- \) = nitrite; \( \text{HNO}_2 \) = nitrous acid

**Common acids to know:** Strong acids: HCl (hydrochloric acid), HBr (hydrobromic acid), HI (hydroiodic acid), HNO\(_3\) (nitric acid), \( \text{H}_2\text{SO}_4 \) (sulfuric acid), HClO\(_4\) (perchloric acid);

Common weak acids: HF (hydrofluoric acid), HCN (hydrocyanic acid), \( \text{H}_3\text{PO}_4 \) (phosphoric acid), \( \text{CH}_3\text{COOH} \) (acetic acid) (often good to memorize acetic acid), \( \text{H}_2\text{CO}_3 \) (carbonic acid)

**COULOMB’S LAW:** Describes the attractive force between a positive (cation) and negative (anion) ion

\[
\text{Force of Attraction} = k \frac{(Q_1)(Q_2)}{d^2}
\]

where \( Q_1, Q_2 \) are ionic charges; \( d \) is distance between the nuclei centers

ionic charges \((Q)\) \(\uparrow\) or ionic radii \(\downarrow\rightarrow d \downarrow \Rightarrow\) Force \(\uparrow \Rightarrow mp/bp \uparrow\) (ionic radii \(\downarrow\) as you go up a column)

Potential Energy = \( U = \text{force x distance} \);

Potential Energy = \( U = k \frac{(Q_1)(Q_2)}{d} \)

**Why do ionic compounds dissolve?** The balance of two energies:

**Lattice Energy:** The energy of attraction between the ions within an ionic compound keeps the ionic compound as a solid.

**Solvation Energy:** The energy between the ions and the solvent molecules (usually water).

ionic solid is soluble (dissolves) when: solvation energy > lattice energy

ionic solid is insoluble (doesn’t dissolve) when: lattice energy > solvation energy

1. Determine the number of protons, neutrons, and electrons present in each of the following atoms/ions.
   a. \( ^{53}_{24}\text{Cr} \)
   b. \( ^{55}_{25}\text{Mn} \)
   c. \( ^{28}_{13}\text{Al}^{3+} \)
   d. \( ^{59}_{28}\text{Ni}^{2+} \)
   e. \( ^{36}_{17}\text{Cl}^- \)

2. Write the complete isotopic symbol with mass and atomic number for each question.
   a. a neutral atom with 28 protons and 30 neutrons
   b. a neutral atom contains 22 protons and 21 neutrons
   c. an oxygen atom with 10 neutrons and the same number of electrons as protons
   d. chromium atom with a mass number of 54 and no charge
   e. a halogen with 35 protons, 35 electrons, and 44 neutrons
   f. an alkali metal with the fewest protons and 4 neutrons and no charge
   g. an iron atom with mass number of 56 and 3 more protons than electrons
   h. sulfide ion with 17 neutrons

3. How many \( p^+, n, \) and \( e^- \) are in the isotopic sulfate (\( \text{SO}_3^{2-} \)) ion: \( ^{33}\text{S}^{16}\text{O}^{18}\text{O}^{18}\text{O}^- \)?

4. Oxygen exists as three isotopes: \( ^{18}\text{O}, \) \( ^{17}\text{O}, \) and \( ^{16}\text{O} \). If the average atomic weight of oxygen is 15.9994, what do you expect the abundance of \( ^{18}\text{O} \) to be approximately?

   <5%  25%  50%  75%  90%
5. An element has two naturally occurring isotopes with the following abundances and masses:

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Atomic Mass (u)</th>
<th>Fractional Abundance</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>84.912</td>
<td>0.7215</td>
</tr>
<tr>
<td>2</td>
<td>86.909</td>
<td>0.2785</td>
</tr>
</tbody>
</table>

What is the average atomic mass of this element? What is the identity of the element?

6. Silver is found commonly as two isotopes: $^{107}$Ag (mass = 106.91u) and $^{109}$Ag (mass = 108.90u). The average atomic mass of silver is 107.869u. What are the percent abundances of the two isotopes, $^{107}$Ag and $^{109}$Ag?

7. There are 3 isotopes of magnesium: $^{24}$Mg (mass = 23.985u), $^{25}$Mg (mass = 24.986u), and $^{26}$Mg (mass = 25.983u). If the abundance of $^{25}$Mg is 10.00% what are the percent abundances of the other two isotopes of magnesium?

8. Which of the following will have the highest melting point? (Hint: ionic radii increase in size as you go down a column in the Periodic Table)
   a. BeO
   b. NH$_3$
   c. MgS
   d. KBr

9. Which of the following compounds would have properties most similar to CaF$_2$?
   a. CO$_2$
   b. SF$_2$
   c. MgCl$_2$
   d. NaCl
   e. CaO

10. Name the following molecules.
    a. NO
    b. PCl$_3$
    c. KBr
    d. Na$_2$CO$_3$
    e. N$_2$O$_4$
    f. K$_2$SO$_4$
    g. HNO$_3$
    h. Ca$_3$(PO$_4$)$_2$
    i. Cu$_2$O
    j. Mn$_3$(PO$_4$)$_2$
    k. Co(CH$_3$COO)$_3$
    l. (NH$_4$)$_2$Cr$_2$O$_7$
    m. H$_3$PO$_4$
    n. CoSO$_4$
    o. HBr
    p. HNO$_2$
    q. KMnO$_4$
    r. CH$_4$

11. Write the chemical formula for the following names.
    a. sulfur trioxide
    b. calcium fluoride
    c. chlorine monobromide
    d. sulfuric acid
    e. sodium sulfate
    f. ammonium nitrate
    g. lithium carbonate
    h. calcium nitrate
    i. manganese(III) sulfate
    j. chromium(VI) oxide
    k. magnesium nitrite
    l. sodium dihydrogen phosphate
    m. titanium(IV) sulfide

12. Which of the following when placed in an ionic compound, will commonly have a charge of −2?
    I. Cl
    II. O
    III. SO$_4$
    IV. Ca
    V. PO$_4$
    VI. NO$_3$
    a. II
    b. II and III
    c. II and IV
    d. II, III, and IV
    e. II and VI

ANSWERS

1. a. 24 p$^+$, 24 e$^-$, 29 n  
   \{bottom number = atomic number = #p$^+$ = 24; top number = mass number = #n + #p$^+$; 53 = #n + 24; #n = 29; #e = #p$^+$ because the isotope is neutral (charge = 0); #e$^-$ = 24\}

2. b. 25 p$^+$, 25 e$^-$, 30 n  
   \{bottom number = atomic number = #p$^+$ = 25; top number = mass number = #n + #p$^+$; 55 = #n + 25; #n = 30; #e = #p$^+$ because the isotope is neutral (charge = 0); #e$^-$ = 25\}

3. c. 13 p$^+$, 10 e$^-$, 15 n  
   \{bottom number = atomic number = #p$^+$ = 13; top number = mass number = #n + #p$^+$; 28 = #n + 13; #n = 15; charge = +3; +3 = #p(+1) + #e(-1); 3 = 13 – x; x = #e = 10\}

4. d. 28 p$^+$, 26 e$^-$, 31 n  
   \{bottom number = atomic number = #p$^+$ = 28; top number = mass number = #n + #p$^+$; 59 = #n + 28; #n = 31; charge = +2; +2 = #p(+1) + #e(-1); 2 = 28 – x; x = #e = 26\}

5. e. 17 p$^+$, 18 e$^-$, 19 n  
   \{bottom number = atomic number = #p$^+$ = 17; top number = mass number = #n + #p$^+$; 36 = #n + 17; #n = 19; charge = -1; -1 = #p(+1) + #e(-1); -1 = 17 – x; x = #e = 18\}
2. 
   a. \( ^{58}_{28}\text{Ni} \) \{28\( ^{+} \) = Ni; bottom number = atomic number = \#p\(^{+}\) = 28; neutral = charge = 0; 
   top number = mass number = \#n + \#p\(^{+}\) = 30 + 28 = 58; \( ^{58}_{28}\text{Ni} \} 
   
   b. \( ^{43}_{22}\text{Ti} \) \{22\( ^{+} \) = Ti; bottom number = atomic number = \#p\(^{+}\) = 22; neutral = charge = 0; 
   top number = mass number = \#n + \#p\(^{+}\) = 21 + 22 = 43; \( ^{43}_{22}\text{Ti} \} 
   
   c. \( ^{18}_{8}\text{O} \) \{O = 8\( ^{+} \) = bottom number = atomic number = \#p\(^{+}\) = 8; neutral = charge = 0; 
   top number = mass number = \#n + \#p\(^{+}\) = 10 + 8 = 18; \( ^{18}_{8}\text{O} \} 
   
   d. \( ^{54}_{24}\text{Cr} \) \{Cr = 24\( ^{+} \) = bottom number = atomic number = \#p\(^{+}\) = 24; neutral = charge = 0; 
   top number = mass number = \#n + \#p\(^{+}\) = 54; \( ^{54}_{24}\text{Cr} \} 
   
   e. \( ^{79}_{35}\text{Br} \) \{35\( ^{+} \) = Br; bottom number = atomic number = \#p\(^{+}\) = 35; \#p\(^{+}\) = \#e\(^{-}\) = charge = 0; 
   top number = mass number = \#n + \#p\(^{+}\) = 44 + 35 = 79; \( ^{79}_{35}\text{Br} \} 
   
   f. \( ^{7}_{3}\text{Li} \) \{alkali metal = Group IA; fewest protons = 3\( ^{+} \) = Li; bottom number = atomic number = \#p\(^{+}\) = 3; \#p\(^{+}\) = \#e\(^{-}\) = charge = 0; 
   top number = mass number = \#n + \#p\(^{+}\) = 4 + 3 = 7; \( ^{7}_{3}\text{Li} \} 
   
   g. \( ^{56}_{26}\text{Fe}^{3+} \) \{iron = Fe = 26\( ^{+} \) = bottom number = atomic number = \#p\(^{+}\) = 26; 3 extra \#p\(^{+}\) than \#e\(^{-}\): 
   charge = \#p\(^{+}\) + \#e\(^{-}\) = 26(+1) + 23(-1) = +3; charge = +3; top number = mass number = \#n + \#p\(^{+}\) = 56; \( ^{56}_{26}\text{Fe}^{3+} \} 
   
   h. \( ^{33}_{16}\text{S}^{2-} \) \{sulfide = S = 16\( ^{+} \) = bottom number = atomic number = \#p\(^{+}\) = 16; ion means charged; Group VIA has a charge of -2 
   (memorized); top number = mass number = \#n + \#p\(^{+}\) = 17 + 16 = 33; \( ^{33}_{16}\text{S}^{2-} \} 

3. protons = 40; neutrons = 45; electrons = 42 \{S: 16\( ^{+} \); O: 8\( ^{+} \) x 3 = 24\( ^{+} \); \#p\(^{+}\) = 16 + 24 = 40\( ^{+} \); 
   \#n: add mass numbers: 33 + 16 + 18 + 18 = 85 = \#n + \#p\(^{+}\); \#n = 85 – 40 = 45; charge = \#p\(^{+}\) + \#e\(^{-}\); -2 = 40 - x; x = 42 = \#e\(^{-}\)\} 

4. <5\% \{since the average atomic weight of O = 15.9994 = 16, then the abundance of \( ^{16}\text{O} \) must be close to 100\% since the 
   average weight is close to 16 and therefore the abundance of \( ^{18}\text{O} \) must be very close to zero otherwise the average atomic mass 
   would be higher than 16\} 

5. 85.47u; Rb \{(84.912)(0.7215) + (86.909)(0.2785) = 85.468u\} 

6. \( ^{107}\text{Ag} \): 51.8\%, and \( ^{109}\text{Ag} \): 48.2\% \{107.869 = 106.91(x) + 108.90(1 - x); x = abundance \( ^{107}\text{Ag} \); 
   107.869 = 106.91x + 108.90 – 108.90x; -1.031 = -1.990x; x = 0.5181 = \( ^{107}\text{Ag} \) abundance; 
   \( ^{109}\text{Ag} \) abundance = 1 – x = 1 – 0.5181 = 0.4819\} 

7. \( ^{24}\text{Mg} \): 78.99\%; \( ^{26}\text{Mg} \): 11.01\% \{x = abundance \( ^{24}\text{Mg} \); y = abundance \( ^{26}\text{Mg} \); sum of fractional abundances: 
   1.000 = x + 0.1000 + y; y = 1.000 – 0.1000 – x; 24.3050 = 23.985(x) + 24.986(0.1000) + 25.983(1 – 0.1000 – x); 
   24.3050 = 23.985x + 2.4986 + 23.3847 – 25.983x; 24.3050 = 25.8833 – 1.998x; -1.5783 = -1.998x; x = 0.7899 = \( ^{24}\text{Mg} \) 
   abundance = 78.99\%; \( ^{26}\text{Mg} \) abundance = 100.00 – 10.00 – 78.99 = 11.01\%\} 

8. a \{NH\(_3\) is a molecular compound (2 nonmetals) and molecular compounds have much lower melting points than ionic 
   compounds; KBr has +1/-1 charges while BeO and MgS have +2/-2 charges; higher charges will yield higher melting points so 
   drop KBr; the other ionic compounds have the same charges: Be\(^{2+}\)O\(^{-2}\), Mg\(^{2+}\)S\(^{-2}\); to differentiate between these the smaller the 
   ions the higher the melting point; ions higher in a column on the Periodic Table are smaller, and therefore since Be\(^{2+}\) is smaller 
   than Mg\(^{2+}\) and likewise O\(^{-2}\) is smaller than S\(^{-2}\), BeO will have the highest melting point\} 

9. c \{elements in the same family/column have similar properties; Mg is in the same column as Ca and Cl is in the same column as 
   F\}
10. a. nitrogen monoxide  b. phosphorus trichloride  c. potassium bromide  d. sodium carbonate  
  e. dinitrogen tetroxide  f. potassium sulfate  g. nitric acid  h. calcium phosphate  
  i. copper(I) oxide  j. manganese(II) phosphate  k. cobalt(III) acetate  l. ammonium dichromate  
  m. phosphoric acid  n. cobalt(II) sulfate  o. hydrobromic acid  p. nitrous acid  q. potassium permanganate  
  r. methane (organic name)

11. a. SO₃  b. CaF₂  c. ClBr  d. H₂SO₄  e. Na₂SO₄  f. NH₄NO₃  g. Li₂CO₃  h. Ca(NO₃)₂  i. Mn₂(SO₄)₃  
  j. CrO₃  k. Mg(NO₂)₂  l. NaH₂PO₄  m. TiS₂

12. b  {memorized}
Polyatomic Ions, Acids, Common Names

(Note: Your instructor will clarify which polyatomic ions are needed for your course; not all of those listed below may need to be memorized!)

- If polyatomic ion name ends with “ate” it has more oxygen atoms than the polyatomic ion whose name ends with “ite” (e.g., nitrate = NO₃⁻ versus nitrite = NO₂⁻)
- If polyatomic ion name ends with “ate” then the acid (adding H⁺ until the ion is neutral) name ends with “ic acid”
  
  e.g., SO₄²⁻ = sulfate → H₂SO₄ = sulfuric acid

- If polyatomic ion name ends with “ite” then the acid (adding H⁺ until the ion is neutral) name ends with “ous acid”
  
  e.g., NO₂⁻ = nitrite → HNO₂ = nitrous acid

- If ion is a single atom (e.g., Cl⁻) then the ion name ends with “ide” (exceptions: CN⁻, cyanide, and OH⁻, hydroxide) and acid name is “hydro+root+ic acid” (e.g., HCl = hydrochloric acid)

NH₄⁺ (ammonium)

H₃O⁺ (hydronium)

OH⁻ (hydroxide)

CH₃COO⁻ = C₂H₃O₂⁻ (acetate)

NO₃⁻ (nitrate)

NO₂⁻ (nitrite)

MnO₄⁻ (permanganate)

CO₃²⁻ (carbonate)

SO₄²⁻ (sulfate)

SO₃²⁻ (sulfite)

ClO₄⁻ (perchlorate)

ClO₃⁻ (chlorate)

ClO₂⁻ (chlorite)

ClO⁻ (hypochlorite)

CN⁻ (cyanide)

CrO₄²⁻ (chromate)

Cr₂O₇²⁻ (dichromate)

PO₄³⁻ (phosphate)

PO₃³⁻ (phosphite)

AsO₄³⁻ (arsenate)

As₂O₃³⁻ (arsenic acid)

F⁻ (fluoride)

Cl⁻ (chloride)

Br⁻ (bromide)

I⁻ (iodide)

CH₃OH = methanol

CH₃CH₂OH = ethanol

C₆H₅COOH = benzoic acid

Some Common Chemicals and Common Names

NH₃ = ammonia

CH₄ = methane

C₂H₆ = CH₃CH₃ = ethane

C₃H₈ = CH₃CH₂CH₃ = propane

H₂O₂ = hydrogen peroxide

C₂H₂O = formic acid