

Periodic Table with Chemistry Formulas

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CHEMISTRY FORMULAS

ATOMS AND SUBATOMIC PARTICLES

A. Atomic number: The number of protons in an atom of a given element. In the periodic table, elements are ordered by increasing atomic number.

Ex: ^{12}C $Z =$ atomic number

B. Mass number: The number of protons plus the number of neutrons in an atom of an element, measured in atomic mass units (amu). 1 amu = mass of one proton or neutron = 1.66×10^{-27} kg.

Ex: ^{12}C $12 =$ mass number

Name	Symbol	Mass	Charge
Proton	p^+	1 amu	$+1$
Neutron	n^0	1 amu	0
Electron	e^-	$1/1837$ amu	-1

C. Isotopes: Various atoms of an element that possess different numbers of neutrons, and therefore different mass numbers.

Ex: Boron has ^{10}B , Boron has ^{11}B

D. Calculating atomic mass: An element's atomic mass number on the periodic table is a weighted average of that element's isotopes.

Ex:

Isotope	Abundance	Mass #	Isotope Mass
Boron-10	20.26%	10	10.0129 amu
Boron-11	79.74%	11	11.0093 amu

Atomic mass = $(20.26)(10.0129 \text{ amu}) + (79.74)(11.0093 \text{ amu})$

MOLECULAR MASS

A. Avogadro's number: 6.022×10^{23}

B. The atomic mass number of an element equals the mass (in grams) of one mole of that element.

Ex: One mole of boron has a mass of 10.81 grams.

MOLECULAR FORMULA AND EMPIRICAL FORMULA

A. Molecular formula: The chemical formula of a compound that gives the actual number of atoms of each element in a molecule of the compound.

Ex: The molecular formula for glucose is $\text{C}_6\text{H}_{12}\text{O}_6$.

B. Empirical formula: A "simplified" expression of a chemical formula.

Ex: The empirical formula for glucose is $\text{C}_2\text{H}_4\text{O}$.

BALANCING CHEMICAL EQUATIONS

A. Check for diatomic molecules early on H_2 and O_2 .

- Balance metals.
- Balance nonmetals.
- Balance oxygen.
- Balance hydrogen.

B. You may change only coefficients, never the formula subscripts! Remember all atoms in the equation have equal numbers on both sides of the equation.

C. Reverse the equation so that coefficients are the smallest whole numbers possible.

PERCENTAGE AND PPM

A. Percentage composition: Tells how much mass any object or element contributes to a compound.

Percentage composition = $\frac{\text{mass of element}}{\text{mass of compound}} \times 100$

B. Parts per million (ppm): In a homogeneous solution, the number of grams of a

particular compound that would be present in a one million gram sample of the compound or solution.

TEMPERATURE CONVERSION

A. Converting between Fahrenheit and Celsius

$$\text{Fahrenheit} = \frac{9}{5}(\text{Celsius}) + 32^\circ\text{F}$$

$$\text{Celsius} = \frac{5}{9}(\text{Fahrenheit} - 32^\circ\text{F})$$

B. Kelvin: The temperature in K, absolute zero, is the theoretical lowest possible temperature, at which all molecular activity ceases.

$$0^\circ\text{K} = -273.15^\circ\text{C} = -459.67^\circ\text{F}$$

$$\text{K} = ^\circ\text{C} + 273.15 \quad ^\circ\text{C} = \text{K} - 273.15$$

IDEAL GAS LAWS

P = pressure
V = volume (expressed in liters)
T = temperature (expressed in Kelvin)
n = number of moles
R = gas constant
Standard temperature and pressure (STP):
 $T = 0^\circ\text{C} = 273.15 \text{ K}$
 $P = 1 \text{ atm} = 760 \text{ mmHg} = 101.3 \text{ kPa}$

A. Boyle's law: $PV = \text{constant}$

$$P_1V_1 = P_2V_2$$

B. Charles' law: $\frac{V}{T} = \text{constant}$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

C. Gay-Lussac's law: $\frac{P}{T} = \text{constant}$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

D. Combined gas law: $\frac{PV}{T} = \text{constant}$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

E. Ideal gas law: $PV = nRT$

$n =$ number of moles, $R =$ ideal gas constant
 $R = 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} = 8.314 \frac{\text{J}}{\text{mol}\cdot\text{K}}$

F. Dalton's law of partial pressures:

$$\frac{P_{\text{gas}}}{P_{\text{total}}} = \frac{n_{\text{gas}}}{n_{\text{total}}}$$

$$\frac{P_{\text{gas}}}{P_{\text{total}}} = \frac{V_{\text{gas}}}{V_{\text{total}}}$$

$$P_{\text{total}} = P_{\text{gas}} + P_{\text{water}}$$

$$P_{\text{water}} = P_{\text{gas}} + P_{\text{water}}$$

$$P_{\text{water}} = P_{\text{gas}} + P_{\text{water}}$$

SOLUBILITY

A. Molarity (M): Measures the number of moles (n) of solute per liter of solution.

$$M = \frac{n}{V}$$

B. Molarity (m): Measures the number of moles (n) of solute dissolved in 1 kg of solvent.

$$m = \frac{n}{W}$$

C. Dilution: The relationship between initial volume (V_i) and final volume (V_f) is inversely proportional to the relationship between initial molarity (M_i) and final molarity (M_f).

$$\frac{V_i}{V_f} = \frac{M_f}{M_i} \quad M_iV_i = M_fV_f$$

COLLIGATIVE PROPERTIES

A. Boiling point: Original pressure. Boiling point when solute is added to pure solvent ΔT_b .

$$\Delta T_b = iK_b \cdot m$$

where $i =$ van't Hoff factor, $K_b =$ boiling point constant of A, $m =$ molality of A, $\Delta T_b =$ the boiling point of A, and $P_0 =$ the vapor pressure of A above pure solvent P_0 .

B. Freezing point depression:

$$\Delta T_f = iK_f \cdot m$$

where $\Delta T_f =$ the amount the freezing point is lowered, $K_f =$ a proportionality constant specific to the solvent, and $m =$ the molality of the solution.

C. Boiling point elevation:

$$\Delta T_b = iK_b \cdot m$$

$\Delta T_b =$ the amount the boiling point is elevated, $K_b =$ a proportionality constant specific to the solvent, and $m =$ the molality of the solution.

D. Osmotic pressure:

$$\pi = iMRT$$

where $\pi =$ the osmotic pressure, $R =$ the gas constant of the solution, $T =$ the total gas constant, and $T =$ the temperature in Kelvin.

CHEMICAL EQUILIBRIA

A. Equilibrium constant expression:

Method using molar concentration: For the reaction $aA + bB \rightleftharpoons cC + dD$

$$K_{eq} = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

Method using gas partial pressures: For the reaction $aA + bB \rightleftharpoons cC + dD$

$$K_{eq} = \frac{(P_C)^c (P_D)^d}{(P_A)^a (P_B)^b}$$

Equilibrium constant $K_{eq} = \frac{[C]^c [D]^d}{[A]^a [B]^b}$

B. Acid-base equilibrium:

1. Acid dissociation reactions: The reaction of a weak acid with H_2O .



Acid dissociation constant: Equilibrium constant for the reaction.

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

The greater the K_a value, the stronger the acid.

2. Base dissociation reactions: The reaction of a weak base with H_2O .



Base dissociation constant: Equilibrium constant for the reaction.

$$K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]}$$

The greater the K_b value, the stronger the base.

3. Dissociation of water: The self-ionization of water.



$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

C. Conjugate acid-base pair:

$$K_a K_b = K_w = 1.0 \times 10^{-14}$$

D. pH and pOH: pH measures $10^{-\text{pH}}$ concentration $10^{-\text{pH}}$.

$$\text{Acids: pH} < 7 \text{ and pOH} > 7$$

$$\text{Bases: pH} > 7 \text{ and pOH} < 7$$

$$\text{Water: pH} = \text{pOH} = 7$$

Calculating pH: $\text{pH} = -\log_{10}[\text{H}_3\text{O}^+]$

Calculating pOH: $\text{pOH} = -\log_{10}[\text{OH}^-]$

For all aqueous solutions:

$$\text{pH} + \text{pOH} = 14$$

concentration with respect to H_2O , assuming the reaction rate is independent of $[\text{H}_2\text{O}]$.

THEORETICAL AND PERCENTAGE YIELD

$$\text{Percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

THERMODYNAMICS

A. Specific heat:

$q = mc\Delta T$
 where $q =$ the heat absorbed by the object, $m =$ the mass of the object, $\Delta T =$ the change in temperature, and $c =$ the specific heat of the substance.

B. Enthalpy (H): In kJ/mol

$$\Delta H = H_{\text{products}} - H_{\text{reactants}}$$

$\Delta H < 0$ for exothermic reactions

$\Delta H > 0$ for endothermic reactions

C. Entropy (S): In J/mol $\cdot$ K

$$\Delta S = S_{\text{products}} - S_{\text{reactants}}$$

D. Gibbs free energy (G): In kJ/mol

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

$\Delta G < 0$ when forward process is spontaneous

$\Delta G > 0$ when reverse process is spontaneous

REDOX REACTIONS AND ELECTROCHEMISTRY

A. Standard reduction potential of a cell:

$$E_{\text{cell}} = E_{\text{cathode}} + E_{\text{anode}}$$

B. Normal equation:

$$E = E^\circ - \frac{RT}{nF} \ln Q$$

where $E =$ the cell potential at any moment, $E^\circ =$ the standard potential, $n =$ the number of electrons in the half-reaction, $R =$ the gas constant, $T =$ the temperature, and $Q =$ the reaction quotient.

C. Standard cell voltage at equilibrium:

$$E_{\text{cell}} = \frac{RT}{nF} \ln K_{eq}$$

where $n =$ the total number of electrons transferred in the redox reaction.

D. Thermodynamics:

$$\Delta G^\circ = -nFE_{\text{cell}}^\circ$$

$$-nFE_{\text{cell}}^\circ = -RT \ln K_{eq}$$

$$E_{\text{cell}}^\circ = \frac{RT}{nF} \ln K_{eq}$$

where $n_0 =$ the starting number of undecayed nuclei, $k =$ the decay constant, and $t =$ the time.

RADIOACTIVE DECAY

A. Particle types:

Alpha	Beta	Gamma	Positron	Neutron
$\frac{4}{2}\text{He}$	$\frac{0}{-1}\text{e}$	γ	$\frac{0}{+1}\text{e}$	$\frac{1}{0}\text{n}$

B. Alpha decay:

$$^{238}_{92}\text{U} \rightarrow ^{234}_{90}\text{Th} + ^{4}_{2}\text{He}$$

C. Beta decay:

$$^{14}_6\text{C} \rightarrow ^{14}_7\text{N} + ^{0}_{-1}\text{e}$$

where $n_0 =$ the starting number of undecayed nuclei, $k =$ the decay constant, and $t =$ the time.

SIGNIFICANT FIGURES

A. Types of significant figures:

— Any nonzero digit is a significant figure.

— Zeros appearing between two nonzero digits are significant.

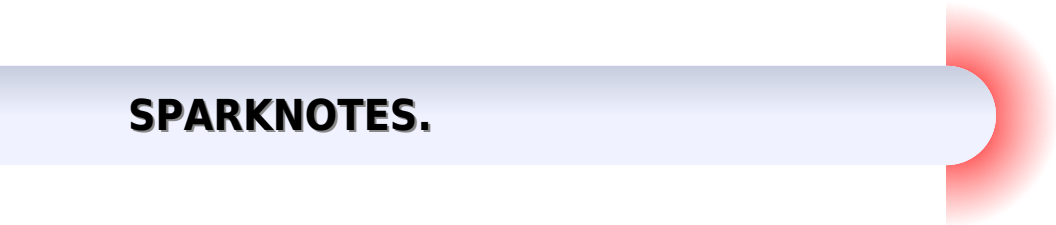
— Zeros appearing at the end of a number, but only if that number contains a decimal point.

B. Multiplication/division: The result of the calculation should have the same number of significant figures as the least precise number being multiplied or divided.

C. Addition/subtraction: The result of the calculation should have the same number of decimal places as the least precise number being added or subtracted.

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