# RIODIC TABI



# 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 ascending atomic number.
  - **Ex**:  ${}^{12}_{6}\mathrm{C}$  6 = 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  $\mathrm{neutron} = 1.66 \times 10^{-27} \mathrm{kg}.$ 
  - Ex:  ${}^{12}_{6}$ 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: Variant atoms of an element that possess different numbers of neutrons. and therefore different mass numbers. Ex: Boron-10: 10B Boron-11: 11B
- D. Calculating atomic mass: An element's atomic mass number on the periodic table is a weighted average of that element's isotopes.

Isotope	Abundance	Mass #	Isotopic Mass
Boron-10	19.78%	10	10.013  amu
Boron-11	80.22%	11	$11.009~\mathrm{amu}$
	(19.78%)(10.01	3.amu) + (80	0.22%)(11.009 amu)

= 10.81 amu

#### **MOLECULAR MASS**

- A. Avogadro's number =  $6.022 \times 10^{23}$  $6.022 \times 10^{23}$  atoms or molecules = 1 mole of that element or substance
- 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

#### 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 C6H12O6.
- B. Empirical formula: A "lowest terms" expression of a chemical formula.
  - Ex: The empirical formula for glucose is  $CH_2O$ .

#### **BALANCING CHEMICAL EQUATIONS**

- 1. Check for diatomic molecules such as H2 and O2.
- 2. Balance metals.
- Balance nonmetals.
- 4. Balance oxygen.
- Balance hydrogen.
- You may change only coefficients, never the formula subscripts! Recount all atoms to be sure you have equal numbers on both sides of the equation.
- 7. Rewrite the equation so that coefficients are the smallest whole numbers possible.

## PERCENTAGE AND PPM

- A. Percentage composition: Tells how much mass an object or element contributes to a compound:
- Percentage composition =  $\frac{\text{mass of element}}{\text{mass of molecule}} \times 100$
- B. Parts per million (ppm): In a compound or solution, the number of grams of a

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

# TEMPERATURE CONVERSION

# A. Converting between Fahrenheit and

$$\begin{split} T_{\text{Fahrenheit}} &= \frac{9^{\circ} \text{F}}{5^{\circ} \text{C}} \left( T_{\text{Celsius}} \right) + 32^{\circ} \text{F} \\ T_{\text{Celsius}} &= \frac{5^{\circ} \text{C}}{9^{\circ} \text{F}} \left( T_{\text{Fahrenheit}} - 32^{\circ} \text{F} \right) \end{split}$$

B. Kelvin: The temperature 0 K, absolute zero, is the theoretical lowest possible temperature, at which all molecular activity ceases.

$$\begin{array}{ll} 0 \; \mathrm{K} = -273.15^{\circ}\mathrm{C} = -459.67^{\circ}\mathrm{F} \\ \mathrm{K} = {}^{\circ}\mathrm{C} + 273.15 \quad {}^{\circ}\mathrm{C} = \mathrm{K} - 273.15 \\ \mathrm{K} = {}^{\circ}\mathrm{F} + 459.67 \quad {}^{\circ}\mathrm{F} = \mathrm{K} - 459.67 \end{array}$$

#### **IDEAL GAS LAWS**

- P = pressure
- V = volume (expressed in liters)
- T = temperature (expressed in Kelvins) Standard temperature and pressure (STP):

 $T = 0^{\circ}C = 273 \text{ K}$ 

$$P=1~\mathrm{atm}=760~\mathrm{mmHg}=101.3~\mathrm{kPa}$$

A. Boyle's law: 
$$PV = constant$$

$$P_1V_1 = P_2V_2$$

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

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

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

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

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

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

#### E. Ideal gas law: PV = nRT

$$n = \text{number of moles}; R = \text{ideal gas constant}$$
  
 $R = 8.31 \frac{\text{L-kPa}}{\text{mol.K}} = 62.4 \frac{\text{L-mmHg}}{\text{mol.K}} = 0.0821 \frac{\text{L-atm}}{\text{mol.K}}$ 

#### F. Dalton's law of partial pressure:

$$\frac{n_{\text{gas 1}}}{n_{\text{total}}} = \frac{P_{\text{gas 1}}}{P_{\text{total}}}$$

$$\frac{n_{\rm gas~2}}{n_{\rm total}} = \frac{P_{\rm gas~2}}{P_{\rm total}}$$

$$P_{\text{total}} = P_{\text{gas 1}} + P_{\text{gas 2}}$$

#### SOLUBILITY

- A. Molarity (M): Measures the number of moles (n) of solute per liters of solution:  $M = \frac{n_{\text{solute}}}{\text{liters of solution}}$
- B. Molality (m): Measures the number of moles (n) of solute dissolved in 1 kg of solvent:

$$m = \frac{n_{\text{solute}}}{\text{kg solvent}}$$

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  $(M_f)$ :

$$\frac{V_f}{V_i} = \frac{M_i}{M_f}$$
  $V_i M_i = V_f M_f$ 

## COLLIGATIVE PROPERTIES

- Raoult's law (vapor-pressure lowering): When solute B is added to pure solvent A,  $P_{\text{actual}} = X_A P_A^{\circ},$ 
  - where  $P_{\text{actual}}$  is the vapor pressure of A after B is added,  $X_A$  is the mole fraction of A, and  $P_{\rm A}^{\circ}$  is the vapor pressure of A above pure solvent A.

#### B. Freezing-point depression:

 $\Delta T_{\epsilon} = K_{\epsilon} m$ . where  $\triangle T_t$  is the amount the freezing point is lowered,  $K_f$  is a proportionality constant specific to the solvent, and m is the molality of the solution.

#### C. Boiling-point elevation:

 $\triangle T_{1} = K_{1}m$ .

 $\triangle T_{\rm b}$  is the amount the boiling point is elevated,  $K_b$  is a proportionality constant specific to the solvent, and m is the molality of the solution.

#### D. Osmotic pressure:

 $\pi = MRT$ .

where  $\pi$  is the osmotic pressure, M is the molarity of the solution, R is the ideal gas constant, and T is the temperature in Kelvins.

#### CHEMICAL EQUILIBRIA

- A. Equilibrium constant expression
- Method using molar concentration:

For the reaction  $aA + bB \rightleftharpoons cC + dD$ :

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

#### Method using gas partial pressures:

For the reaction

$$CO_2(g) + H_2(g) \rightleftharpoons CO(g) + H_2O(\ell)$$
:

Equilibrium constant:  $K_{\mathrm{eq}} = \frac{P_{\mathrm{CO}}}{P_{\mathrm{CO}_2} P_{\mathrm{H}_2}}$ 

#### B. Acid-Base equilibrium

1. Acid dissociation reactions: The reaction of a weak acid with H2O:

 $HA(aq) + H_2O \rightleftharpoons H_3O^+(aq) + A^-(aq)$ Acid dissociation constant (equilibrium constant for the reaction):

$$K_{\rm a} = \frac{[{
m H_3O^+}][{
m A}^-]}{[{
m HA}]}$$

The greater the  $K_{\rm a}$  value, the stronger the acid.

2. Base dissociation reactions: The reaction of a weak base with H2O:

 $B(aq) + H_2O \rightleftharpoons BH^+(aq) + OH^-(aq)$ Base dissociation constant (equilibrium constant for the reaction):

$$K_{\rm b} = \frac{[\rm BH^+][\rm OH^-]}{[\rm B]}$$

The greater the  $K_{\rm b}$  value, the stronger the

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

$$2H_2O(\ell) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$
  
 $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ 

4. Conjugate acid-base pair:

$$K_{\rm a}K_{\rm b} = K_{\rm w} = 1.0 \times 10^{-14}$$

- 5. pH and pOH: pH measures [H+]; pOH measures [OH-].
  - Acids: pH < 7 and pOH > 7
  - Bases: pH > 7 and pOH < 7
  - Water: pH = pOH = 7

Calculating pH:  $pH = -\log_{10}[H_3O^+]$ Calculating  $pOH: pH = -\log_{10}[OH^-]$ For all aqueous solutions:

[OH<sup>-</sup>][H<sub>3</sub>O<sup>+</sup>] = 
$$1 \times 10^{-14}$$
  
 $pH + pOH = 14$ 

## REACTION RATES

- For the reaction  $aA + bB \longrightarrow cC + dD$ : Rate =  $k[A]^m[B]^n$ , where k = rate constant; m = rate order with respect to A; n = rateorder with respect to B.
- Ex: If the rate law is  $k[A]^2[B]$ , the reaction is second order with respect to A and first order with respect to B.

**Ex:** If the rate law is  $k[A]^0 = k$ , the reaction is zero order with respect to A, meaning the reaction rate is independent of [A].

# THEORETICAL AND PERCENTAGE YIELD

Percentage yield =  $100 \times \frac{\text{Actual yield}}{\text{Theoretical yield}}$ 

## THERMODYNAMICS

#### A. Specific heat:

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

- B. Enthalpy (H): in kJ  $\mathrm{mol}^{-1}$ 
  - $\triangle H = H_{\text{product}} H_{\text{reactants}}$
  - $\triangle H < 0$  for exothermic reactions
  - $\triangle H > 0$  for endothermic reactions
- C. Entropy (5): in  $J \text{ mol}^{-1} K^{-1}$

$$\triangle S = S_{\text{product}} - S_{\text{reactants}}$$

- D. Gibbs free energy (G):  $in kJ mol^{-1}$  $\triangle G = \triangle H - T \triangle S$ 
  - $\triangle G = 0$  when at equilibrium
  - $\triangle G < 0$  when forward process is
  - spontaneous  $\triangle G > 0$  when reverse process is
  - spontaneous

#### REDOX REACTIONS AND ELECTROCHEMISTRY

- A. Standard reduction potential of a cell:
- $E_{\text{cell}}^{\circ} = E_{\text{reduction}} + E_{\text{oxidation}}$

## B. Nernst equation:

 $E=E_{\mathrm{cell}}^{\diamond}-\frac{RT}{nF}\ln Q=\frac{.0592~\mathrm{V}}{n}\log_{10}Q$ where R is the universal gas law constant, T is in Kelvins. n is the number of electrons in the half-reaction, F (Faraday's constant) =  $96485 \text{ C mol}^{-1}$ , and Q is the reaction quotient.

C. Standard cell voltage at equilibrium:  $E_{\text{cell}}^{\circ} = \frac{RT}{nF} \ln K_{\text{eq}} = \frac{.0592 \text{ V}}{n} \log_{10} K_{\text{eq}}$ where n is the total number of electrons

# transferred in the redox reaction

D. Thermodynamics: 
$$\triangle G^{\circ} = -nFE_{\mathrm{cell}}^{\circ} \\ -nFE_{\mathrm{cell}}^{\circ} = -RT\ln K_{\mathrm{eq}}$$

# RADIOACTIVE DECAY

A. Particle types Alpha Beta Gamma Positron Neutron

#### $_{-1}^{0}e$ $_{0}^{0}\gamma$ Ex: Alpha decay:

 $^{238}_{92}U \longrightarrow ^{4}_{2}He + ^{234}_{90}Th + 2^{0}_{0}\gamma$ 

B. Exponential decay:

 $n = n_0 e^{-\lambda t}$ where  $n_0$  is the starting number of undecayed nuclei,  $\lambda$  is the decay constant. and t is the time.

# SIGNIFICANT FIGURES

# A. Types of significant figures

- Any nonzero digits in a number
- Zeros appearing between two nonzero digits in a number
- Zeros appearing at the end of a number, but only if that number contains a decimal point
- 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.

CADATON STATES WAS STREET SHE HOLD   1.3   TOWN C NUMBER     ELCTRON STATES   TOWN C	18   18   18   18   18   18   18   18	18	EN 26 EN 26 FIP I2.8 AR 124 Xenon [Kri 4d <sup>10</sup> 55 <sup>-5</sup> 59 <sup>6</sup>	EN —	Ununoctium	3 27 EN 127 EA -50 EA -50 EA 172 AR 172 AR 172 LUteflum [Xe] 4ft*5d/6s²	3 103   Lawrencium   Rnl 5f*6d/7s²		
1.3   A   A   A   A   A   A   A   A   A	HALOGENS    7 VIIB   1 1 1 1 1 1 1 1 1 1 1 1 1 1 1 1 1 1 1	3.5.7 17 3.5.7 17 3.5.7 17 3.5.7 17 3.5.7 17 3.5.7 17 3.5.7 17 3.5 3.5.7 17 3.5 3.5 3.5 3.5 3.5 3.5 3.5 3.5 3.5 3.5	1266 1266 1266 1266 133 133 126,9047 <b>2616</b> 1344 <sup>10</sup> 55 <sup>2</sup> 5p <sup>3</sup>	EN 220 FR -270 ASTATINE   Astatine   Xel 4f*5d"6s*6p*		3 70 N 11 Pb 625 R 1.94 N 173.04 Werbium	2, 3 EN 13 EN 6-5 FIP 6-65 AR - (259) Nobelium RNI 5f <sup>47</sup> 75 <sup>2</sup>		
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